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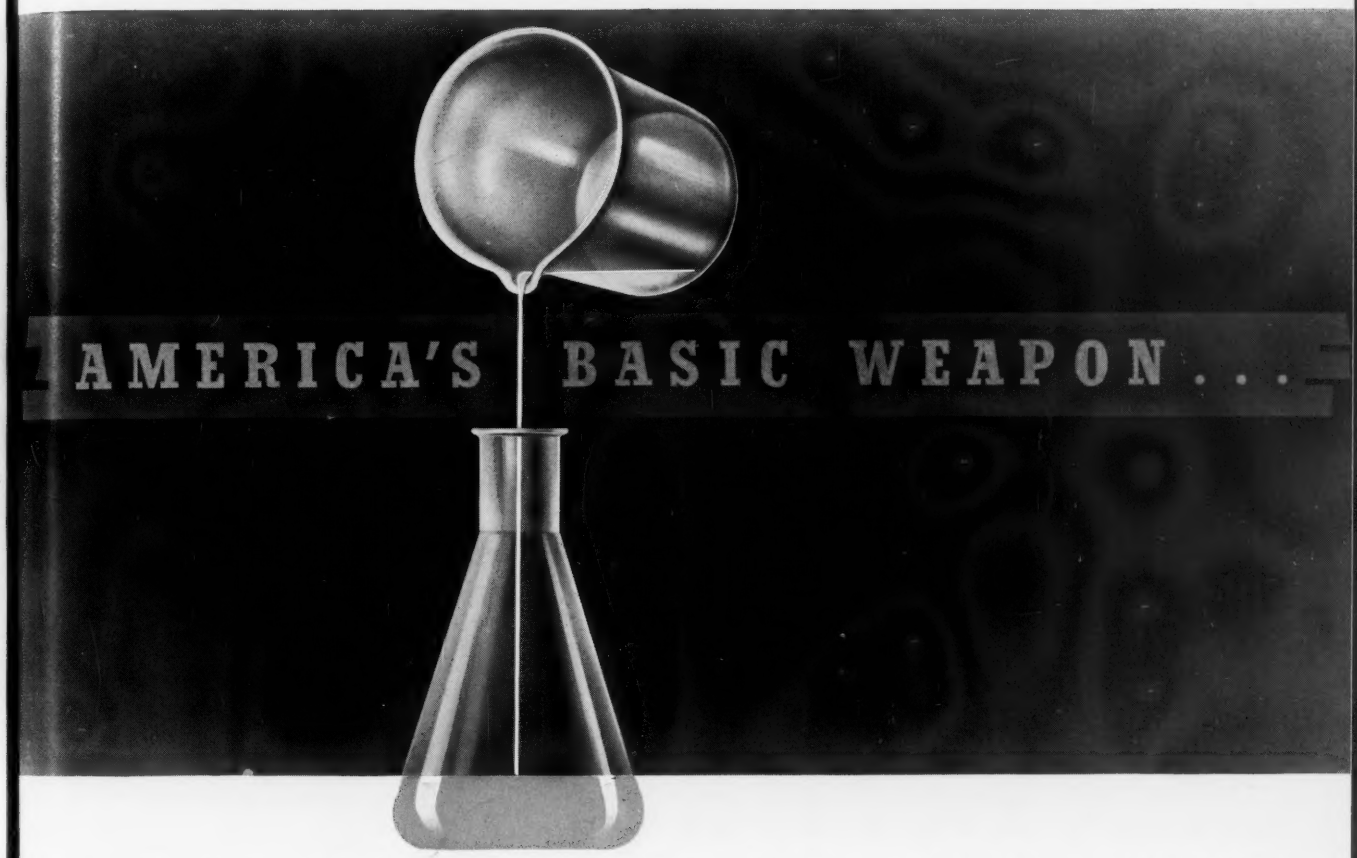
CHEMICAL EDUCATION



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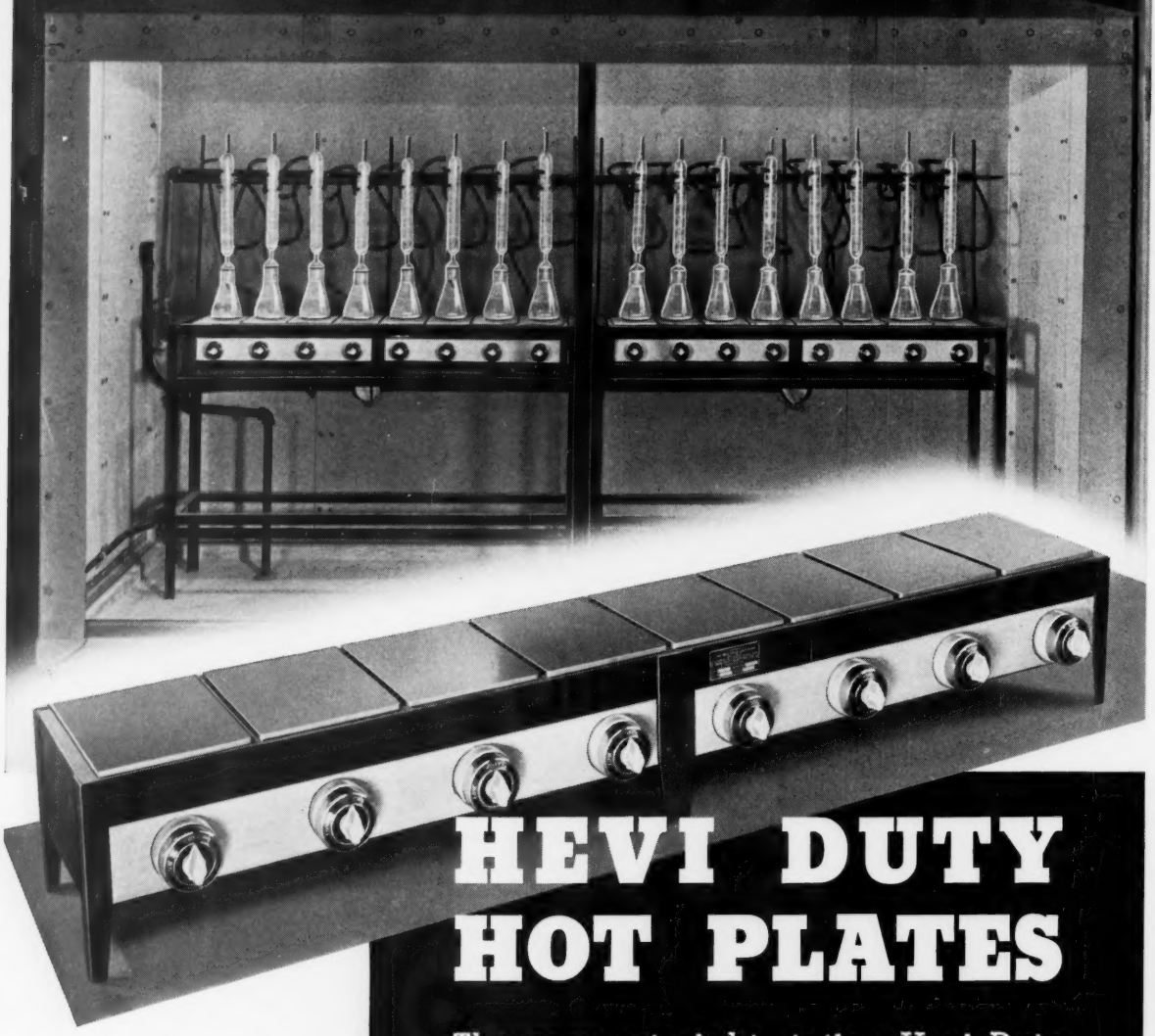

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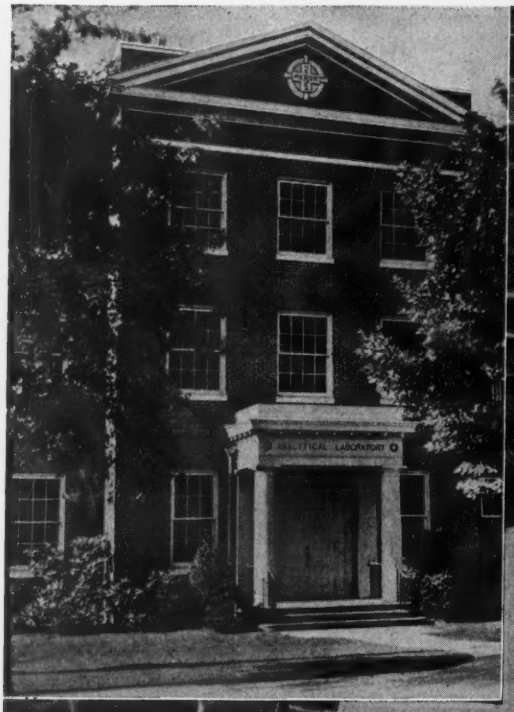

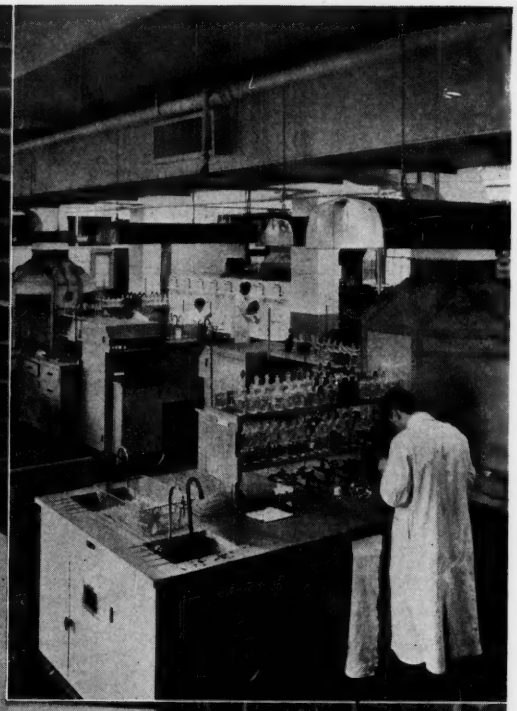

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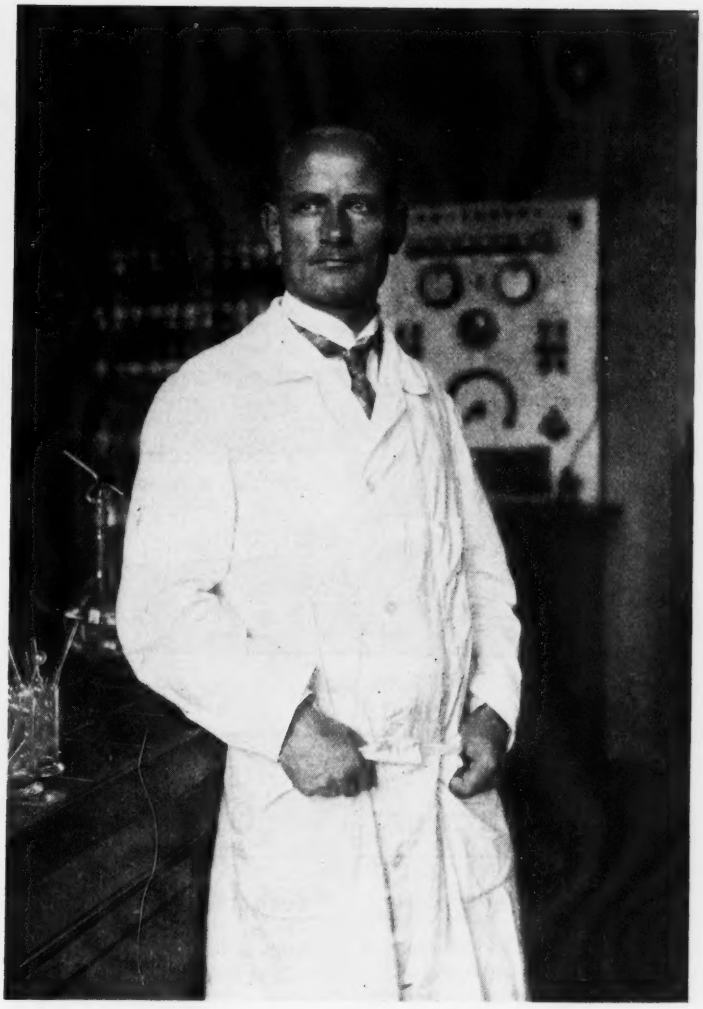

Otto Ruff (1871–1939)

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(For further information about Otto Ruff, see page 496.)

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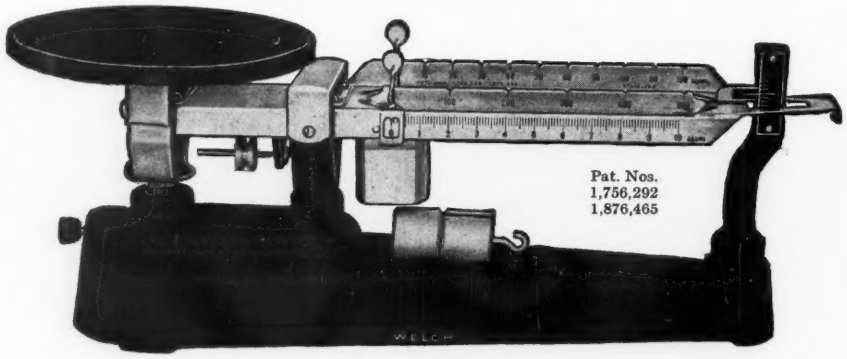

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Editor's Outlook

SO! WE'RE GOING into high gear. Fine! And as far as most of us are concerned there is no need of waiting until after the elections. Nor will we object if our whole educational system feels the strain pretty quickly.

At the recent Institute on Education and the War, Commissioner Paul V. McNutt said: "There is no excuse for any young man or woman to be in college preparing for any profession not directly useful to the war effort." Coming, as this statement did, from a man who is at the same time the head of the War Manpower Commission and the Overlord of the Bureau of Education, it was a bombshell to higher educators, but one for which their air-raid warning system should have prepared them. Followed up—and the rest of the discussion at the Institute followed it up very effectively—it means a drastic overhauling of the curricula of our colleges and universities, and, to a lesser extent, our secondary schools. Accordingly, our "liberal arts" programs become a luxury which, for the duration, we must largely do without—along with aluminum teakettles, new cars, and Sunday joyrides.

However, it is one thing to lay down a policy, and quite another to follow it consistently and sensibly. And when General Somervell says: "Every able-bodied boy is destined at the appointed age for the armed services," we recognize the old Army weakness, resulting from the premise that all men are alike except in point of age and term of military service. We shudder to think where this point of view will lead us if it permeates *all* of our Army's operations.

This Ethiopian has been lurking in the woodpile for some time, with vague murmurs revealing his presence; now he has been chased out into the open. If we may take this statement at its face value it means that we are not going to train any more scientific and technical brains for the vital production industry. For you have to take brains where you find them, and we flatly disagree that all we need can be found in the incapacitated and the women (God bless 'em). This, of course, completely nullifies the spirit of the Selective Service Act, which we always *thought* was intended to be selective.

We have before us a news release from the Office of War Information which says: "Business establishments engaged in war production are going to be told within the next few weeks where and how they can obtain chemists, engineers, metallurgists, and other professional and scientifically trained men, Paul V. McNutt, Chairman of the War Manpower Commission announced today." We certainly hope this information is passed along soon; before long business establishments engaged in war production may not be *able* to obtain chemists, engineers, and metallurgists. Nor, perhaps, will industry or civilians be able to find a doctor.

And here's another OWI release quoting Donald Nelson, Chairman of the War Production Board, as saying: "The battle of production is not yet won." And still another, saying that a single General Grant tank needs about 26 tons of steel, almost 600 pounds of copper, over 500 pounds of chromium, and more than 600 pounds of manganese, aluminum, lead, and zinc, and urging that somehow these metals be forthcoming. The mail also brought us some dazzling posters, with the stunning caption in four-inch letters: MORE PRODUCTION. Also the "Official Plan Book" explaining the War Production Drive.

How do all these parts of the jig-saw puzzle go together? Can't we ever manage to coordinate our demands for man power?

Production is not a matter of labor alone, but brains as well. It is scarcely necessary to repeat the voluminous evidence of the need of technical training in industry, and the essential position of chemists, among others. Much is being done, of course, in the matter of re-training workers for war industry. You can take a man from a doughnut factory and in a short time make him an acceptable welder or riveter. But not within some years can you make the same man into a chemist capable of developing, operating, or even controlling an industrial process. In other words, the production army —like the combat army—must have its officers as well as its enlisted men. And the officers' training schools for the production army are in our universities and technical schools. If the Army prevents us from carrying on this vital branch of training, it will have to be responsible for the consequences two or three years from now. We contend that the brains and capacity necessary for this most important branch of service are out of all proportion to anything that the Army needs for its own personnel, and that such capacity should have a triple-A priority (or whatever the highest category happens to be at the moment) in the competition for ability.

The numbers involved are entirely trivial in comparison to their importance. The Army talks vaguely of a force of ten million men. All that the production army asks for its young officers in training is a few thousands—of the right ones. A pile of service cards or a column of statistics will not do; *this* army is interested in individuals.

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What's Been Going On

COMPLEX maps and ocean navigation charts which can be read in total darkness, military control panels lighted by invisible rays are now practical for war use through the development of a new group of ultra-violet black light sources.

All of the new fluorescent type of black light devices, ranging from a walnut-sized bulb to a four-foot glass tube, use a newly discovered chemical coating which transforms short ray ultra-violet into near ultra-violet or black light.

Maps which must be read under black-out conditions either on land or sea can actually be made to light up in the dark in several different colors for different types of information. Charts coated with phosphorescent chemicals act like storage batteries of light, absorbing a quantity of illumination in less than a minute, and then releasing it in the dark over a period of several hours.

Trail blazing with fluorescent powders or paints is another one of the important possible new uses for black light. By this method, markings left on trees, stones, and bushes remain invisible until picked out in the darkness by ultra-violet spotlights.

The importance of the use of luminescence in its application to war and black-out problems is being brought out repeatedly. From the laboratories of the Calco Chemical Company comes the following statement:

"Never has this subject been of more interest to the public than under the present emergency. Of importance is the fact that few people realize that there are three distinct classifications in the luminescent field. The first, radioactivity, is a phenomenon which is inherent in certain matters. The main war use for this type of material is for instrument dials where cost is a secondary consideration. Phosphorescence is a property which can be applied to some substances and gives them the ability to glow in the dark after exposure to sunlight, ultra-violet light, or incandescent light. Fluorescence, on the other hand, is a property which can be applied to practically any material by painting, coating, or dyeing, and requires a source of ultra-violet light in order to activate it. Fluorescent materials do not have a useful after-glow.

"From the same laboratories that produced fluorescent carpet, now used in many moving picture theaters, have come many more developments along the same lines. It is interesting to note that among the most common uses of fluorescent markings is that of dials, knobs, and panels for use in air fighter work where the ordinary incandescent lamp on an instrument board cuts down the visibility of the pilot, but ultra-violet activated fluorescent dial markings leave the operator with a clear vision. Where it is possible to expose 'marked' objects to sunlight or other types of natural illumination, it has been found that phosphorescent coatings have a wide application, such as marking field or other equipment which must be handled or assembled under black-out conditions. Thus, if repairs are to be made on transport units, many of them can be done by coating the tools with phosphorescent paint, and work can continue in the dark. In this group it is interesting to note that phosphorescent material can be applied by the popular decalcomania method.

"While many people have been under the impression that fluorescence is less handy to use than other forms of luminescence, it is interesting to note that many forms of activating lamps are now available which require as little as $2^1/_2$ volts to operate, and will show objects within a working radius. As a matter of fact, even with transportation units, a 6-volt head lamp of the ultra-violet type is all that is needed for the purpose of activating road markings. So extensive has become the use of luminescent materials that one will find in many dime stores packages of luminescent identification markers on sale for a few cents; or one can purchase at the local paint store enough luminescent paint to apply on a large barn.

"A danger, however, not generally realized is that a mass of areas of luminescence has high visibility, the same as does ordinary light, and caution should be exercised in using large quantities of these materials out-of-doors. However, even this disadvantage can be capitalized, as it is possible to paint the interior of a room with phosphorescent paint, and, in case of failure of power sources, sufficient light will be available therefrom to carry on ordinary activities. Fluorescent tapes can also be prepared from wool, rayon, nylon, cotton, acetate, or other textile fibers for use in outlining doorways, entrances, etc., during black-outs. These can be activated by a minimum of black light, for example, by the use of a $2^1/_2$-watt Argon bulb in an ordinary 110-volt socket. Special ultra-violet hand lamps and 6-volt storage battery lamps are available for activating fluorescent objects, either during power failures or in places where normal power facilities are not available."

NEW COURSES AT THE UNIVERSITY OF PITTSBURGH

THE UNIVERSITY of Pittsburgh will offer a course in "Military Chemistry and Chemical Agents," each trimester beginning September 28. It will be based on "Technical Manual 3-215" of the War Department and will cover all of the fundamentals and principles which are required by the Manual. Two fully illustrated lectures will be given weekly by Dr. A. L. Robinson. The purpose of the course is to train prospective registrants for military service.

The Chemistry Department will also offer free evening courses for training in chemistry for war industries. The purpose of these courses is to prepare new men and women to replace those who have entered the Armed Forces and to train individuals already employed in industry for more advanced chemical service. The courses planned include: Introductory and Advanced Metallurgical Analysis; Microscopical Testing of Industrial Materials; Electrical and Optical Instruments for Analytical Control; Food and Water Inspection; Practical Vitamin Assaying; Industrial Quantitative Organic Analysis.

An elementary course on Chemical Fundamentals for War Workers is also planned as preparation for those listed above.

A descriptive bulletin of the courses may be obtained by writing to Dr. Herbert Spencer, Regional Coordinator of ESMWT, Pennsylvania College for Women, Pittsburgh, Pennsylvania. Fall courses are scheduled to begin October 12, and registration will take place during the preceding week.

Lecture Demonstrations

JAMES H. WALTON *University of Wisconsin, Madison, Wisconsin*

EVERY LECTURER in general chemistry has developed new demonstrations or improved some of the old ones. With the feeling that a more frequent interchange of ideas would be of great benefit to teachers of chemistry, the writer has been led to describe certain new and modified demonstrations and techniques which may prove helpful.

Hydrogen-Chlorine Explosions. The explosion of a mixture of these gases can be strikingly demonstrated as follows: Use a 1 by 8-inch glass test tube, fill it with a saturated brine solution, and invert the filled tube in a pan of brine. Displace about half the brine in the tube with chlorine gas and the rest with hydrogen gas. Stopper the tube with a *paraffined* cork stopper, *C*. This is very important, inasmuch as chlorine attacks ordinary cork and is used up after a few minutes' contact. Place the test tube inside a metal tube arranged as shown in Figure 1. The tube can, of course, be held in a clamp attached to a ring stand instead of being mounted as shown in the figure. *B* is a hole in the metal tube through which a beam of light can be focused on the gas mixture. The writer uses a small carbon arc lantern for this purpose. This results in an explosion which blows the cork stopper about 60 feet without breaking the tube. That the longer wave lengths of light are ineffective can be shown by placing a sheet of red glass between *B* and the source of light.

When first tried, the violence of the recoil resulted in breaking many of the test tubes. This was prevented by placing a thin pad of soft rubber, *A*, in the breech of the gun.

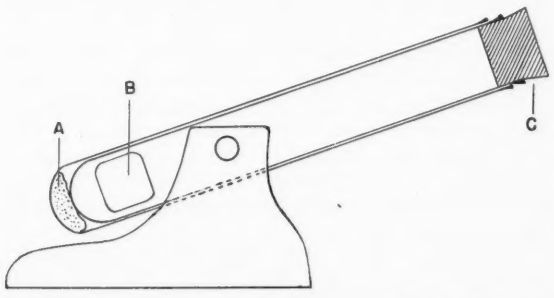


FIGURE 1.—THE EXPLOSION OF HYDROGEN AND CHLORINE

In discussing this reaction it is desirable to demonstrate the effect of heat on the mixture of the two gases. This can be done by touching the glass tube at *B* with a tiny flame. The writer uses the flame from a small blast lamp for this purpose. A flashlight powder or burning magnesium ribbon can be used as sources of light energy, but they are not entirely satisfactory because the intense light detracts attention from the main objective of the demonstration.

Oxy-Hydrogen Mixture. The two gases are exploded in the iron tube made from ordinary water pipe and capped as shown in Figure 2. The source of heat for ignition is an ordinary spark plug, *P*. The tube is filled with the proper volumes of the two gases by using a float, *S*, which is a piece of cork stopper carrying a piece of thin wire the length of the tube. This is used as follows: fill the iron tube with water, invert it in a rather deep pan of water, then place the float, cork up, inside the iron tube. The lower tip of the wire can be seen below the end of the pipe. Fill the tube with hydrogen gas until two-thirds of the wire of the float can be seen, then bubble in oxygen gas until the tube is full. Stopper tightly, using a cork stopper. When current from a coil connected with the spark plug is turned on, the resulting explosion will blow the stopper about 80 feet.

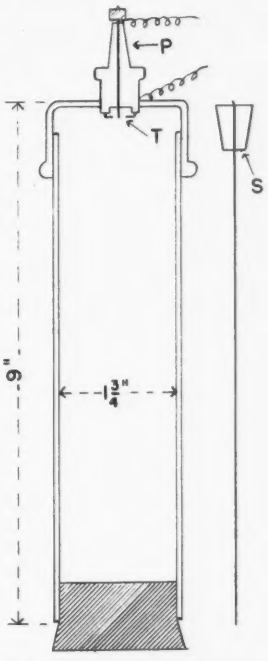


FIGURE 2.—THE EXPLOSION OF HYDROGEN AND OXYGEN

Attention is called to the fact that the explosion may be prevented by water adhering to the terminals of the spark plug. This can usually be removed by sharply rapping the tube of gas two or three times against the edge of the desk. This tube can also be used to demonstrate the explosion of carbon monoxide and oxygen.

The Reduction of Cupric Oxide. A familiar demonstration of oxidation-reduction in the lecture room is the interaction of hydrogen and hot copper oxide. In the experience of the writer this is best shown by heating a brightly polished copper crucible in the air and then allowing the hot blackened crucible to come in contact with hydrogen gas. The demonstration is not always satisfactory unless certain details of construction and manipulation of apparatus are carefully followed.

The apparatus used is shown in Figure 3. The crucible, *C*, is about 4 inches in diameter at the open end.

It is supported on a disc of transite which carries a $^1/_4$-inch layer of sand held in place by the rings of transite which are fastened to the base. The sand is used to seal the bottoms of the crucible and the funnel, *F*. The funnel is raised and the crucible is heated with a Meker burner until the surface is black but not hot enough to glow. Without turning off the burner, the 8-inch glass funnel is then lowered and hydrogen admitted slowly into the stem. This results in the reduction of the copper oxide and the restoration of the original copper color.

Raising the funnel brings back the black color. It is not necessary to shut off the hydrogen gas while the copper is being oxidized again. The oxidation and reduction can be carried on intermittently. The change of color is very striking and if carried out as described can easily be seen by large classes. If the crucible is heated to too high a temperature, the hydrogen usually ignites with a slight but usually harmless explosion, which, however, diverts the attention of the class from the real object of the demonstration. Sealing the crucible and the funnel with sand keeps the hydrogen away from the flame. A little practice will show the conditions necessary for a successful demonstration.

If a copper crucible cannot be purchased, a satisfactory substitute, *D*, can be made from sheet copper, using high melting silver solder.

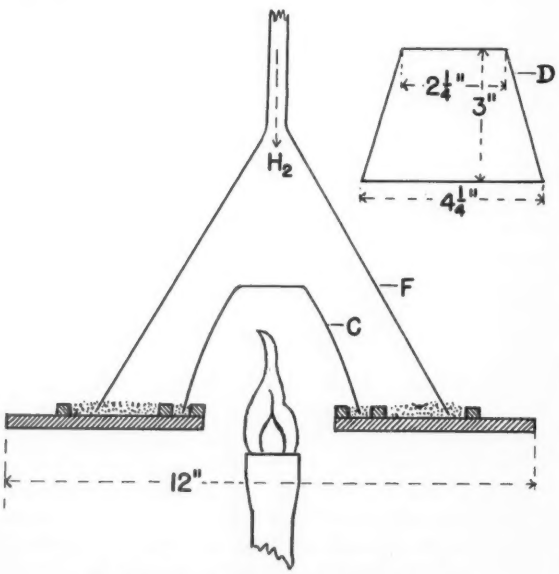


FIGURE 3.—REDUCTION OF CUPRIC OXIDE

Partial Miscibility. Shake ethyl ether and water thoroughly and allow the layers of liquid to separate in the tube shown in Figure 4. Remove some of the ether layer by means of the upper stopcock and show the presence of water by stirring the liquid with a few grams of anhydrous copper sulfate. Remove some of the water layer and place it in a small casserole. It will contain enough ether to permit the vapors to be ignited with a match. To make the two layers in the tube more apparent, a crystal of iodine may be dissolved in the ether.

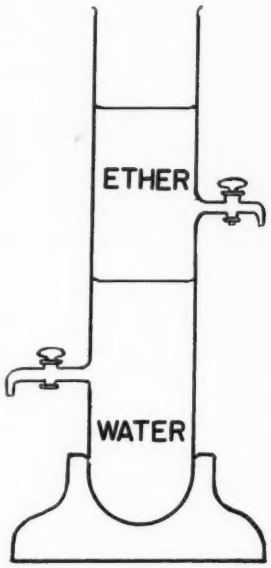


FIGURE 4.—PARTIAL MISCIBILITY

Specific Heat. In spite of the fact that the air thermometer is familiar to everyone, it is not in general used as a means of qualitatively demonstrating temperature changes to large classes. Figure 5 gives the dimensions of the one used by the writer. The stem is made from standard 6-mm. glass tubing. Any dark aqueous solution may be used for the liquid. The opening *A* can be provided with a glass stopcock or closed with a bit of rubber tubing and a piece of glass rod. The thermometer has been used in the following demonstrations:

Use 2-liter beakers each filled with water and provided with an air thermometer. Mark the height of the liquid in each thermometer by means of a strip of paper. Heat 500-g. pieces of zinc, iron, etc., in an oil bath to about 250°C. Each metal should have a piece of copper wire attached so that it can be easily handled. Remove metals from the oil bath, place one in each beaker, and stir the water.

The specific heats of two liquids such as water and turpentine are compared by heating two 500-g. brass weights in the oil bath and introducing one into each beaker.

Oxidation of Aluminum. Place a 10 by 10-inch piece of thin aluminum foil on the lecture table. Using a piece of cotton, wet the foil thoroughly on both sides with a saturated mercuric chloride solution and wrap the foil around the bulb of an air thermometer. A great increase in temperature will be observed, attended by the usual formation of white aluminum oxide on the surface of the foil.

Solubility of Gases. The tube shown in Figure 6 is very useful for showing the solubility of gases. It is best filled with mercury by connecting the top with the suction pump, which saves mercury and the temper of the lecture assistant. The mercury can then be displaced with the desired gas in the usual manner. The solubility of ammonia in water and in activated charcoal is familiar to everyone, but the writer has not run across any mention of the interaction of ammonia and copper sulfate. This is carried out by making a 3 to 4-g. pellet of anhydrous copper sulfate and introducing it into a tube of ammonia gas. To show the solubility

of carbon dioxide in alkali, fill the cup at the top of the tube with sodium hydroxide solution and allow the liquid to run slowly into a tube of carbon dioxide gas. This gives good surface contact between liquid and gas and the absorption takes place rapidly.

Burning Hydrogen. The apparatus shown in Figure 7 is used to burn hydrogen and condense and collect the water formed. The stopcock at the bottom of the apparatus permits the liquid to be drawn off as desired. A slight suction must be applied at the tube *A*.

CIVILIAN DEFENSE

Gas Defense. In discussing the principles on which the gas mask is based, a demonstration of adsorption of a gas, ammonia by charcoal, for example, will interest the class. Selective adsorption can be shown by the apparatus illustrated in Figure 8. Tube *C* contains absorbent cotton moistened with benzene. Close the stopcock in *A*, open the cock in *B*, and pass a stream of hydrogen gas through the apparatus. The enriched gas will burn at *D* with a luminous flame. Close the cock in *B* and allow the gas to pass through *A*, which is filled with loosely packed silica gel or activated carbon. The benzene vapors are adsorbed and the gas now burns with a non-luminous flame.

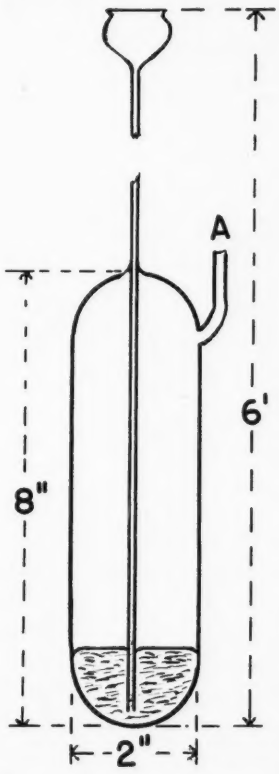


FIGURE 5.—AIR THERMOMETER

Vapors and Heavy Gases. The properties of heavy vapors and dense gases such as chlorine should be discussed and, if possible, illustrated experimentally. For vapors use the demonstration described later under "Inflammable Vapors"; for gases, any one of the numerous experiments that show how carbon dioxide gas flows and seeks the lowest level.

Incendiaries. Thermite in a crucible or other suitable container can be readily ignited as follows: mix fresh sodium peroxide and aluminum powder by placing a teaspoonful of each in a tin can and rotating the can. Pour the mixture on the thermite and mix some of it with the thermite. Allow three or four drops of water to fall on the ignition mixture. In a few seconds a vigorous reaction will take place, attended by the ignition of the thermite.

WARNING! This ignition mixture should be kept away from sweaty hands; further, do not use sodium peroxide that has been standing in the air.

Oxidation by Bleaching Powder. Inasmuch as bleaching powder is so important in decontamination, a demonstration of its powerful oxidizing properties is of interest. For this purpose use H.T.H. or "Perchloron." Place about a teaspoonful on a tile so as to form a small crater and allow a few drops of glycerine to fall into the depression. A vigorous reaction takes place, resulting in the ignition of the glycerine.

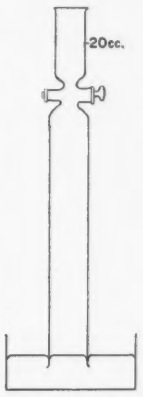


FIGURE 6.—SOLUBILITY OF GASES

ACCESSORIES

The lecturer in general chemistry can learn much from the methods the theatrical producer uses to focus the attention of the audience on the performance of a singer or dancer. All other individuals go into the background, and the star is continually in the spotlight. Many demonstrations may be dramatized by the same technic: a prominent and isolated spot on the lecture table and good illumination. Force rtain demonstrations the luminous background illustrated in Figure 9 is very useful. This consists of a box, *A*, about 2 feet long and 1 foot high which has a ground glass front. Inside the box are two 100-watt bulbs. *B* in the figure is a turntable, something like the "Lazy Susan" once placed in the center of the dinner table to make the condiments available to the family. This carries the illuminated background and enables the lecturer to turn the table so that the students who sit at the side of the lecture room do not always have to look at a piece of apparatus from an acute angle.

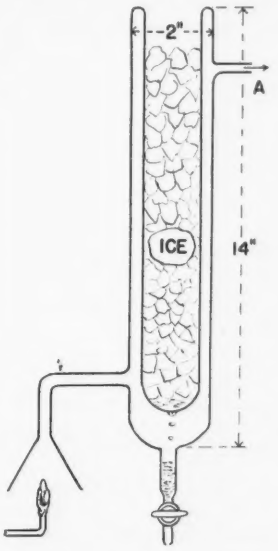


FIGURE 7.—WATER FROM BURNING HYDROGEN

Illumination from overhead may be helped by placing beakers, etc., directly under an ordinary desk lamp.

A more extensive overhead lighting system is shown in Figure 10, in which the illumination is provided by a number of photoflood bulbs carried in *B*, which is built to shield the class from glare. The background *C* is a reversible 12-inch board, black on one side and white

on the other. The top *B* is counter-balanced by a lead weight, *A*, at each end so that it can be tipped back if it is necessary to use the blackboard behind it. Clamps and binding posts can be attached to *C* if desirable.

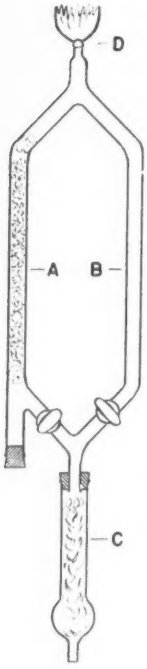


FIGURE 8.—SELECTIVE ADSORPTION

For the removal of fumes while working with chlorine or similar gases, the hood shown in Figure 11 has proved very satisfactory. *A* is the upper half of a Pyrex flask 11 inches in diameter. It is cemented to the metal pipe which can be inserted in the socket *B* as needed. *B*, which extends into the basement of the laboratory, carries the motor, *M*, of a discarded household vacuum cleaner, the speed of which can be regulated by a proper resistance. The exhaust is connected with one of the laboratory flues. The motor is not entirely noiseless, but it is turned on for short periods only. The present motor has been in use for about six years, but has not been seriously affected by corroding gases.

Inflammable Vapors. The insidious nature of inflammable vapors is convincingly shown by allowing them to run down a trough and come in contact with a flame. Arrange a piece of eaves trough about 16 feet long and 4 inches wide so that one end rests on the floor and the other is elevated about 6 feet. Place a lighted candle at the lower end. In the upper end of the trough place a piece of cheesecloth that has been wet with about 20 cc. of high test gasoline or ethyl ether. The cloth should be wet, but no liquid should drip from it. In less than a minute the vapors will be ignited by the candle flame and the flame will run up the trough. The piece of cheesecloth should be removed before it takes fire. This demonstration is more striking if carried out in a darkened room.

The number of casualties which result every year from attempts to dry-clean with gasoline in a warm basement where there is a gas-heated hot water tank make this demonstration of general educational value.

Liquid Sulfur Dioxide. This liquid, which is a nonconductor of electricity, gives a good conducting solution when a few crystals of potassium iodide are dissolved in it. As a consequence, it can be used to advantage in the discussion of electrolytic dissociation. When the solution of potassium iodide in sulfur dioxide is allowed to evaporate, red crystals of the solvated compound are formed ($KI \cdot 4SO_2$). Upon standing, this compound turns white as a result of the loss of SO_2. These properties can be demonstrated in connection with the discussion of the properties of solutions.

The liquefaction of SO_2, the determination of the boiling point of the liquid, the conductivity of the solution, and a preparation of the solvated potassium iodide make an interesting and comparatively simple experiment for some of the more advanced students of beginning chemistry. Liquefaction of SO_2 is best carried out with dry ice.

Photosynthesis by Plants. The demonstration of this important reaction involves the availability of the proper plant and a suitable source of light. *Elodea*, a water plant obtained from the Department of Botany of the university, is very satisfactory when exposed to the powerful light from photoflood bulbs. The demonstration is carried out by placing a few sprigs of the plant in a 2-liter beaker which is then filled with water

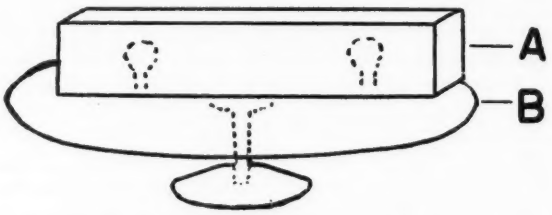


FIGURE 9.—"LAZY SUSAN" AND LUMINOUS BACKGROUND

that is practically saturated with CO_2. A funnel is then inverted over the elodea. Bubbles of oxygen released on the leaves pass up through the stem of the funnel and can be collected by displacement of water in a test tube inverted over the stem of the funnel. Two photoflood bulbs, screened from the audience, are used. They are placed about 8 inches from the side of the beaker. With this source of light, a stream of small

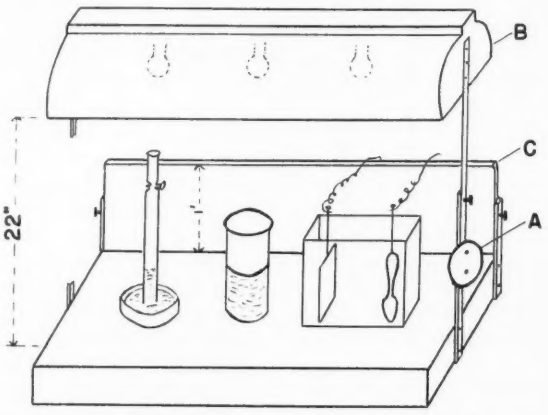


FIGURE 10.—ILLUMINATION OF LECTURE DEMONSTRATIONS

bubbles of oxygen is given off for a period of several hours.

MISCELLANEOUS

In the discussion of the properties of solutions, the decrease in solubility with increase in temperature can be illustrated by the use of calcium butyrate. This salt is easily prepared by standard methods. A tube

containing 100 cc. of a saturated water solution at room temperature is immersed in a beaker of water heated to near the boiling point. A white precipitate forms in two or three minutes. If the tube is then removed and immersed in ice water, the calcium butyrate goes back into solution.

The addition of turpentine to nitric acid to show the oxidizing power of the latter usually but not always results in the ignition of the turpentine. This can be made more certain if 2 to 3 cc. of turpentine are cautiously added to a 50–50 mixture of concentrated sulfuric and nitric acids.

A simple method to illustrate the displacement of iodine or bromine by chlorine is to soak a piece of blotting paper with the halide salt and drop it into a jar of chlorine gas.

To add powdered antimony to chlorine gas use a salt shaker or an ordinary kitchen flour shaker.

To show the enormous stresses set up by the sudden cooling of glass, "Baloney" flasks, available from some of the chemical supply companies, are more satisfactory than Prince Rupert's drops. These flasks are about four inches long and can easily be seen by the class. When a fragment of carborundum, pinhead size, is dropped into one of these flasks, the flask immediately breaks into bits. This can be made more mystifying as follows: Rap the flask smartly against the desk top to show its strength, hold the flask horizontally and introduce the end of a 12-inch piece of ordinary glass tubing, which contains a piece of carborundum, into the neck of the flask. Blow the carborundum into the flask. With a large class, the flask is more visible if it has been colored by dipping it in ink.

Professor Ernst Cohen used to relate how he won a dinner bet by lighting a lump of sugar. The trick, is accomplished by first rubbing the sugar in tobacco ash. This demonstration has frequently been used by the writer when lecturing on catalysis. The reaction is said to be catalyzed by the metallic oxides in the tobacco ash.

Monoclinic Sulfur. The formation of monoclinic sulfur crystals described by Fowles[1] has been modified as follows: Dissolve 200 grams of rhombic sulfur by heating with 1400 cc. of toluene. A 3-liter round-bottomed flask is convenient. Place the flask of hot sulfur solution in a box to insure slow cooling and insulate the flask well with cotton batting. In a few hours long needles of sulfur are formed, which later become covered with small crystals of rhombic shape. The two crystal forms can be distinctly seen.

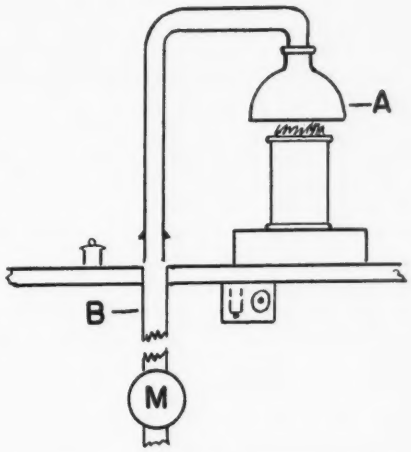


FIGURE 11.—HOOD FOR LECTURE TABLE

The construction of the apparatus described has been made possible by the help of Mr. James B. Davis, university glass blower, and Mr. Lee K. Henke, chief mechanician for the chemistry department.

[1] FOWLES, "Lecture experiments in chemistry," The Blakiston Company, Philadelphia, **1937**, p. 125.

CIVIL SERVICE

INCREASING numbers of scientifically and technically trained men and women will be required for the war effort this year and next. Junior Metallurgists are urgently needed to conduct investigative, developmental, or production work in various branches of metallurgy; to assist in the design, construction, installation, and operation of metallurgical equipment; or to perform metallurgical work in the recovery or fabrication of metals.

The U. S. Civil Service Commission is recruiting Junior Metallurgists under a new Announcement (No. 254) for which the qualifications are: (1) completion of a four-year college course in metallurgy or metallurgical engineering, or (2) completion of a four-year course in chemistry, geology, physics, or engineering, *supplemented* by (*a*) one year of paid experience in metallurgy (college teaching in metallurgy is acceptable), or (*b*) 15 semester hours in metallurgy or metallurgical engineering, or (*c*) completion of two War Training Courses in metallurgy.

There is provision for the acceptance of applications from college seniors or graduate students who expect to complete the required courses within six months after filing applications.

Announcements and application forms may be obtained at any first- or second-class post office, or from the Civil Service Commission, Washington, D. C.

To remove calcimine from smooth plaster, if the film is too thick to come off readily by washing, soak the surface thoroughly with a solution of trisodium phosphate and water. While the surface is wet, scrape it with a 3- or 4-inch broadknife. If calcimine is sufficiently soaked it will remove in ribbons.

—LAUCKS

Project Teaching in High-School Chemistry

II. Models and Demonstration Apparatus

WILLIAM C. CURTIS *Dunbar High School, Washington, D. C.*

THE CONSTRUCTION of models and demonstration apparatus develops both the academic and mechanical skills of superior students in high-school chemistry. Such projects offer an excellent opportunity for students to plan and carry out their own purposeful and self-directed activities. Moreover, they tend to satisfy the urge to create and enable students to correlate the knowledge and skills acquired in other subjects with their work in chemistry.

The student should first select a chemical principle, process, or concept or an important application to modern life. A model or demonstration apparatus should then be planned to illustrate or demonstrate the principle or idea selected.[1] The student should gather or collect as much of the material needed in the completion of the project as possible. When completed it should be neat, attractive, and dynamic.[2]

Carefully planned inspection trips to such museums as the New York or Chicago Museum of Science and Industry and the National Academy of Sciences, Washington, D. C., will serve to familiarize the students with examples of skilfully constructed and protected scientific exhibits. Close-up pictures of chemical plants and laboratories and their surroundings are of tremendous value to students undertaking the construction of models.

Some examples of suitable models and types of demonstration apparatus which high-school chemistry teachers may appropriately lead their students to "discover" are given below. Many of these should prove of value in connection with our defense efforts. Some of them can be used for instructional purposes. Others may be useful in working out local camouflage and blackout problems.

MODELS

(1) Atomic and molecular structures. *Cf., Sci. Leaflet,* **13,** 117 (1939).
(2) Bomb resistant shelters.
(3) Bromine extraction plant.
(4) Coal mine. *Cf., ibid.,* **13,** 6 (1939).
(5) Coal tar plant. *Cf.,* J. CHEM. EDUC., **5,** 1112 (1928).
(6) Cyclotron.
(7) Dioramas. *Cf., ibid.,* **13,** 412 (1936); *ibid.,* **17,** 397 (1940).
  (*a*) Famous chemists and their laboratories.
  (*b*) Early and modern chemical laboratories.
(8) Frasch process. *Cf., ibid.,* **8,** 630 (1931); *ibid.,* **10,** 309 (1933).
(9) Gas plant.
(10) Gasproof shelters (ventilated and non-ventilated types).
(11) Glass factory.
(12) High explosive bombs.
(13) Hydrochloric acid plant.*
(14) Hydrogenation of vegetable oils.
(15) Iodine production plant.
(16) Incendiary bombs.
(17) Liquid air plant.
(18) Liquid ammonia plant. *Cf., ibid.,* **8,** 2062 (1931).
(19) Liquid carbon dioxide plant.
(20) Lime kiln.
(21) Manufacture of steel.*
  (*a*) Bessemer process.
  (*b*) Open hearth process.
(22) Manufacture of sulfuric acid.
  (*a*) Contact process. *Cf., ibid.,* **8,** 2239 (1931).
  (*b*) Lead chamber process. *Cf., ibid.,* **18,** 47 (1941).
(23) Manufacture of rayon.
(24) Molecular still.
(25) Nitric acid plant.
(26) Paper making machine (Fourdrinier).
(27) Pasteurization plant.
(28) Petroleum refinery.
(29) Refrigerating machine.
(30) Respiratory calorimeter.
(31) Soap kettle.
(32) Solvay process.*
(33) Sewage disposal plant. *Cf., ibid.,* **17,** 167 (1940).
(34) Synthetic gasoline plant.
(35) Synthetic rubber plant.
(36) Typical oil field.
(37) Unexploded bombs (model showing how to detect the presence of unexploded bombs).
(38) Water purification plant.
  (*a*) Aeration plant.
  (*b*) Chlorination plant.
  (*c*) Filtration plant.*
(39) Water works plant.

DEMONSTRATION APPARATUS

(1) Adsorption. *Cf., ibid.,* **18,** 432 (1941).
(2) Aluminum extraction. *Cf., ibid.,* **18,** 193 (1941).
(3) Alpha-ray track apparatus.*
(4) Balance. *Cf., ibid.,* **10,** 118 (1933); *ibid.,* **17,** 227 (1940).
(5) Conductivity of metals.*
(6) Conductivity of solutions and fused salts.* *Cf., ibid.,* **17,** 381 (1940).
(7) Colloid cataphoresis. *Cf., ibid.,* **6,** 334 (1929); *Sch. Sci. Math.,* **36,** 502 (1936).
(8) Constant temperature oven.
(9) Cottrell process of smoke precipitation.* *Cf.,* J. CHEM. EDUC., **5,** 1001 (1928); *Sci. Leaflet,* **14,** 812 (1941).
(10) Density of liquids.*
(11) Destructive distillation of coal and wood.
(12) Dry cell. *Cf.,* J. CHEM. EDUC., **6,** 778 (1929).
(13) Electric arc.
(14) Electric furnace (resistance, induction, or arc type).
(15) Electric glass cutter.
(16) Electrolysis of sodium chloride (Nelson or Vorce cells).
(17) Electroplating.*

[1] SHAW, *American Physics Teacher,* **7,** 165 (June, 1939).
[2] *Cf.,* CURTIS, J. CHEM. EDUC., **18,** 293 (1941).

* References on marked items are given in J. CHEM. EDUC., **17,** 389 (1940).

(18) Fireless cooker.
(19) Fixation of nitrogen.* *Cf., ibid.*, **10,** 348, 498 (1933).
(20) Gas mask. *Cf., ibid.*, **19,** 322 (1942).
(21) General behavior of gases.
  (*a*) Boyle's law. *Cf., ibid.*, **18,** 15 (1941).
  (*b*) Charles' law. *Cf., ibid.*, **15,** 394 (1941).
  (*c*) Laws of partial pressures and diffusion. *Cf., ibid.*, **12,** 355 (1935).
(22) Hot plate.*
(23) Hydrogen sulfide generator (odorless).*
(24) Hydrometer.
(25) Hygrometer.
(26) Magnesium extraction.
(27) Oudin coil. *Cf., Science Observer*, **2,** 16 (1940).
(28) Ozonizer. *Cf.*, J. CHEM. EDUC., **5,** 1493 (1928).
(29) Peltier effect.
(30) Photoelectric cell.*
(31) Polariscope. *Cf., Science Observer*, **2,** 6 (1940); *Chem. Leaflet*, **15,** 1020 (1942).
(32) Potentiometer.
(33) Spectroscope (prism or grating type). *Cf., Sch. Sci. Math.*, **41,** 346 (1941).
(34) Sensitive flame.
(35) Telescope.*
(36) Telsa coil.*
(37) Tyndall effect.
(38) Ultramicroscope. *Cf.*, J. CHEM. EDUC., **18,** 378 (1941).
(39) Ultraviolet lamp.
(40) Vacuum pump.
(41) Voltaic cell.

SCENERY SUGGESTIONS

Groundwork—built on corrugated cardboard with wooden frame.
1. Irregular surface—scraps of cardboard with tapered edges or cloth painted with wax.
2. Earth—plaster or plastic paint molded over a base of wire screening or burlap and painted ground color with various shades of brown, water-soluble, inside wall paint or sifted earth on glue or paint. A damp or wet effect may be obtained by using the latter method and applying shellac after the glue or paint has dried.
3. Grass—sawdust dyed in two or three shades of green with artist's oil colors and glued to groundwork, hemp, dried grasses, thin strips of paper.
4. Shrubs and trees—small dried plants such as dyed Norwegian lichen, gypsophila (baby's breath), steel wool sprayed with glue and sprinkled with olive green sawdust, sponge rubber, etc.
5. Snow—paraffin wax.
6. Water surfaces—ripple glass, plain window glass, old mirrors, crumpled foil from cigaret packages, thin or heavy sheets of celluloid, paraffin wax colored with oil paint.

Buildings—wood, corrugated cardboard, cardboard or congoleum cork painted with water color.
Figures—armature made of twisted wire, Plasteline (permanent modeling clay), or pipe cleaners (for small figures) painted with thick water colors.
Novel lighting effects—Prismatic film.

SOURCES OF CONSTRUCTION MATERIALS

DuPont Plastacele (an excellent protective material)—Arthur & Bro., 67 West 44th Street, New York; Maxwell Display Corporation, 240 West 40th Street, New York.
Electric Motors, Mechanisms, etc.—Merkle-Korff Gear Co., 211 North Morgan Street, Chicago, Illinois; Speedway Manufacturing Co., 1839 South 52nd Avenue, Cicero, Illinois.
Lumber Supplies—T. A. Foley Lumber Co., Paris, Illinois.
Metal Goods—Metal Goods Corporation, 5249 Brown Ave., St. Louis, Missouri.
Modeling Clay—Western Stoneware Co., Manmouth, Illinois.
Plastic and Water-Soluble Paints—Sherwin-Williams, Graphic Arts and Display Division, Cleveland, Ohio.
Prismatic Film—Optaz Manufacturing Co., 206 Broadway, New York.

A NOTE ON THE LABORATORY DETERMINATION OF THE HEAT OF TRANSITION OF SULFUR

PAUL PORTER SUTTON

North Carolina State College of Agriculture and Engineering, Raleigh, North Carolina

ACCORDING to the First Law of Thermodynamics, the heat of transition from rhombic to monoclinic sulfur is exactly equal to the difference between the heats of solution of each species in a given solvent, provided they are so measured that the resulting mixtures have in both cases the same composition. Since the dissolving of this substance in carbon tetrachloride is endothermic, it is, in principle, possible to determine these heats of solution in this liquid with an ordinary vacuum type student calorimeter, such being the procedure recommended by Daniels, Mathews, and Williams.[1] Two factors, however, render such measurements unsatisfactory unless special equipment is available: first, the low solubility of the element in this medium, which makes for a small temperature change; and second, the fact that the solvent is so volatile that its rapid evaporation causes a greater temperature drop than the effect to be observed.

In seeking for a modification of this experiment, the writer was able to eliminate both of the above-mentioned objections by replacing the carbon tetrachloride with a carbon disulfide-water mixture. The water covers the carbon disulfide, largely preventing evaporation, without influencing its solvent properties. Using a Beckmann thermometer, it was found that once temperature equilibrium was reached the calorimeter could then be handled in the usual manner.[2] With 350 cc. of carbon disulfide and 50 cc. of water, 10-gram samples of rhombic and monoclinic sulfur gave temperature depressions of the order of 1.4 and 1.3 degrees, respectively. These are easily measured, and a final value of the molar heat of transition within 10 per cent of the accepted can readily be obtained.

[1] DANIELS, MATHEWS, AND WILLIAMS, "Experimental physical chemistry," McGraw-Hill Book Company, New York, **1934,** p. 93.

[2] *Ibid.*, pp. 91–2.

Dyes and Dyeing

HANOR A. WEBB *George Peabody College for Teachers, Nashville, Tennessee*

NATURE is colorful. Even before men lived upon the earth the sky was blue, the forests green, the rocks gray, the soil brown. Sunrises and sunsets splashed golds and reds across the horizon, and from every summer shower the rainbows arched.

No one knows whether the animals enjoy nature's colors, but they appreciate their own finery. The gay feathers of male birds, so useful in their courtships, are examples.

When the Human Family finally spread over the earth, its members were very conscious of color. In primitive languages there are few separate names for color, but they are attached to objects: blue is "sky color"; yellow, "fire color"; green, "leaf color"; brown, "earth color"; and so on. But for bright red—which is rare in nature—some expressive words such as "the amazing color" were invented.

From earliest times the Human Family has colored its clothing for ornament. Egyptian mummies, placed in their tombs six thousand years ago, were wrapped in linen dyed blue with indigo. In the familiar Bible story (Genesis 37:3) the boy Joseph received "a coat of many colours" from his father. The walls of Pompeii's ancient ruins are decorated with gay pictures proving that Greek and Roman costumes were ornamented with many colored stripes and figures. Kings and nobles of the Mediterranean countries wore woolen robes colored with Tyrian purple, a juice squeezed from a tiny shellfish, *Murex*, and so costly that the dyed wool was worth the equivalent of a hundred dollars a pound!

We of today, as color-minded children of our ancestors, make the most lavish use of color in human history. Not only are our clothes brilliant, but our dwellings are bright with color inside and out, our automobiles are the world's gayest vehicles. Our magazine covers, our signs, our wrapping paper, avoid mere black and white. Merchants sell more articles if they are colored, regardless of the quality. Even the lips and fingernails of our ladies are dyed.

We use color so freely today because coloring materials are abundant and cheap. For this the chemists deserve most thanks.

We shall now tell of chemical achievements in the invention of dyes from coal tar, and chemical control of dyeing. As a background for the colorful present we shall tell brief stories of the past.

THE MINERAL COLORS

The colored earths—clays and minerals—gave early men their first hues for adornment of skin and garments. Centuries later painters of the Middle Ages purified and blended these minerals most artistically. Their favorite pigments were the *ochres*, quartz minerals mixed with oxides of iron, giving yellows to browns; the *siennas*, containing oxides of manganese; the *umbers*, with both iron and manganese oxides. From iron oxide itself they prepared paints ranging from red to black. Lead oxides gave yellows to red, chromium ores supplied the greens, soot the black.

When chemists began to study the minerals they invented several new paints, such as chrome yellow (lead chromate), zinc yellow (zinc chromate), cadmium yellow (cadmium sulfide), iron blue (ferric ferrocyanide), chrome green (chromium oxide), ultramarine blue (aluminum silicate, sodium sulfide), lead white (lead carbonate-acetate), zinc white (zinc oxide), titanium white (titanium oxide), and many others. These mineral pigments are much used today in paints and printing inks, but little in dyeing cloth.

THE VEGETABLE DYES

Early men learned by accident that crushed plants would color their skins and garments. As time went by, the boiled barks, seeds, stems, shoots, leaves, berries, roots of countless plants became dyestuffs of commerce. Famous were indigo for blues, birch leaves for yellow, madder for red—but there were hundreds of others. From the tropics, ships brought loads of dyewoods, chiefly the heartwoods of forest trees from which the dyes could be boiled out.

Today dyewoods are still shipped to this country in rather large quantities from the West Indies, Central, and South America. Most important is logwood for blues, grays, and blacks; fustic for yellows; brazilwood for reds. Cutch from India supplies tans and browns. Not only cloth, but leather, furs, and bristles are colored in this country by the tropical dyewood extracts. Metal salts called mordants (explained on later pages) must be used to get strong, durable colors.

THE ANIMAL DYES

The most expensive dye in history was obtained 2000 years ago by fishermen on the Mediterranean Sea, who gathered baskets of a tiny mollusk, *Murex*, from the shallow waters. After pulling the slimy little body from its shell, a few drops of juice were squeezed into a vessel. After fermentation, and with secret methods, mordants, and washings, a bluish red color was given to wool. So many dyers with their "tricks of the trade" lived at Tyre on the Mediterranean's eastern coast that the color was called "Tyrian purple." The labor and the secrecy increased the cost of the dyed wool over a thousand times, and only the Emperors of Rome

and other rulers could wear these "royal purple" garments. No Tyrian purple is made today.

Another humble creature, the scale-insect *Coccus cacti* of Mexico, has served for centuries as the source of the brightest scarlet dye for wool. The dried bodies are ground to powder, from which a dye called cochineal is boiled out. Compounds of tin and oxalic acid are used with the cochineal in the famous "British red-coat formula." Cochineal is still on the market in small amounts for dyeing wool, but it will not color cotton.

THE COAL TAR COLORS

A truly colorful civilization began in the laboratory of a London college during the Easter holiday of 1856. One young man, eighteen-year-old William Henry Perkin, made the discovery. As a schoolboy of thirteen he had missed his lunch between classes to visit the college down the street, and listen to the lectures of a noted German chemist, A. W. von Hoffmann. Observing the boy's interest, Professor Hoffmann gave him laboratory tasks to do, and soon put him to work on problems of chemical investigation. The professor and his young helper were especially anxious to make quinine artificially.

On the particular day of the discovery, when all other students were taking vacations, young Perkin was trying to add oxygen to aniline oil—a chemical made from benzene, distilled in turn from coal tar. He was using potassium bichromate, an orange-colored solution of great oxidizing power. He expected white crystals, but instead got a black, sticky mass. Disgusted, he soaked this tarry stuff in alcohol in order to clean his glass flask, and noticed that the alcohol turned purple.

Less thoughtful persons would have poured tar and alcohol into the waste jars. Instead, young Perkin hurried to a nearby store, and bought about a yard of white silk ribbon, which he dipped into the alcohol. The color the silk acquired would not wash out—it was truly dyed to a rich, rather reddish purple.

If Perkin was an observant young chemist, Hoffmann, his teacher, had set him a fine example. Now all the skill of the older, experienced man was needed—for when the experiment was repeated with new aniline the color did not come! Hoffmann decided that the old aniline used by Perkin contained a small amount of some impurity. Much testing was done to find this, and only a chemist of Hoffmann's knowledge and experience could have separated and identified the toluidine (methyl-amino-benzene) that was present, in but a few drops to the quart, in the aniline (amino-benzene) Perkin had used on his lucky day. But, really, did luck play much part in the discovery of the first coal-tar dye? Rather was it not faithfulness and attention to details on the part of the young chemist, and experience on the part of the old chemist, that brought success?

Young Perkin soon proved himself a good business man, too. With Professor Hoffmann's chemical advice and his father's financial aid, he built a factory to manufacture this first man-made color. He called it *mauve*, a French word for a purple color from the mallow plant. He persuaded enough dyers of silk and wool in London to try his color to make it popular with well-dressed people. He experimented in his laboratory with other colors with some success. He made a modest fortune, was honored by knighthood, and lived to welcome delegates from all the world to the Golden Jubilee of 1906, the fiftieth anniversary of his discovery. There he cut into small pieces the very ribbon he had first dyed with a coal-tar color, and gave the small but priceless samples to the most noted chemists who were present. He died the next year, in 1907, aged 69.

Having begun in London, it would seem that the making of dyes from coal tar would have become one of England's greatest industries. But Perkin remained almost the only chemist in the country who studied and manufactured dyes. The Germans, however, rapidly trained chemists who discovered new dyes at the rate of dozens each year. Huge dye factories were built in Germany from which dyes were shipped to all the world. German chemists beat Perkin by a single day in filing patents for a way to make alizarin—the madder-root red—from coal-tar chemicals. At a cost of $5,000,000 and fifteen years of experimenting, the German Baeyer made coal-tar indigo. The cultivation of madder and indigo plants practically ceased, for the artificial products were both better and cheaper than the natural dyes.

The near-monopoly or one-nation control of coal-tar dyes by Germany continued until the First World War. Then the small dye industries of Great Britain and the United States were greatly enlarged. After the war, the dye manufacturers of the United States were protected from German competition by high tariffs—charges against foreign products brought into our ports. Factories in the United States produced 120,190,000 pounds of dyes in 1939, with a value of over $70,000,000.

HOW DYES ARE MADE

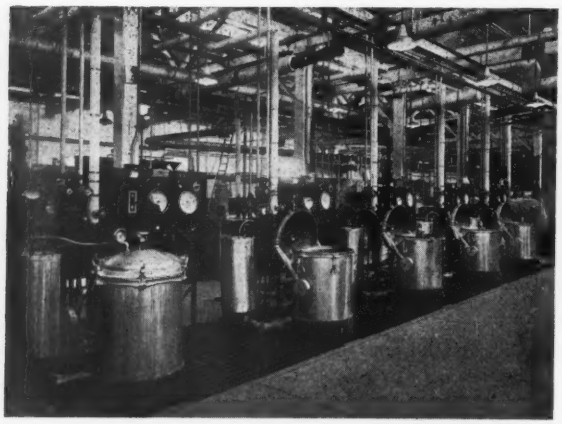
In the early days the name *aniline dyes* was applied to these artificial colors, since the first one was made from aniline oil. New dyes were invented, however, that were no relatives of aniline, but all came from coal-tar chemicals; the name *coal-tar dyes* has therefore had wide use. Within recent years a more accurate name, *synthetic organic colors*, has been adopted by most chemists. The word *synthetic* means "put together" in Greek, and makes a distinction between the synthetic or man-made dyes and the natural dyes of nature's manufacture. The word *organic*[1] implies former life, in this case the plants that formed the coal, and makes a distinction between organic dyes from ancient living matter and inorganic dyes from non-living minerals.

[1] It should be noted, however, that the chemical term *organic* no longer has any real relation to life or substances produced by living organisms, but in general refers merely to *compounds of carbon*.

Since so many dyes are not related to aniline, and there are really no dyes as such in coal tar, the recent name—synthetic organic colors—is best.

The substances in synthetic dyestuffs have passed through three principal stages—as crudes, intermediates, and dyes.

The Crudes. When coal is heated in closed ovens, coal gas and coal tar are driven off while coke remains in the oven. When coal tar itself is heated in retorts, some score of "crudes" are boiled off, which may be separated later by distilling, since their boiling points vary. When purified, these products, still called "crudes" though crude no longer, are colorless liquids or white crystals. The ten selected crudes that are the starting points of most dyes bear the typically musical names of benzene, toluene, xylene, phenol,



Courtesy Taylor Instrument Company, Rochester, New York

A DYE FACTORY

cresol, naphthalene, anthracene, methyl anthracene, phenanthrene, and carbozol. All contain the "benzene ring," all contain the chemical elements carbon and hydrogen, some also contain oxygen. Each has an odor which the trained nose detects, usually with pleasure.

The crudes most familiar to you are probably phenol, commonly called carbolic acid, and naphthalene, known to you as "moth balls."

Up to 1914 most of the coal tar produced in coke and gas ovens of the United States was used merely for street paving and roofs. This nation's limited needs for crudes was purchased from Germany.

The Intermediates. From the score of coal-tar crudes some 500 "intermediates" are manufactured, each by its own processes of combination with a host of chemicals, such as sulfuric, hydrochloric, and nitric acids, ammonia, alcohol, chlorine, hydrogen, and many others. You may find several of these intermediates in pure form on the shelves of any school chemistry laboratory, especially if courses in organic chemistry are offered. Aniline, picric acid, benzoic acid, salicylic acid, nitrobenzene, are some of the labels you should look for. A few of the intermediates—such as picric acid—are themselves dyestuffs of poor quality.

These intermediates are used to make many other things besides dyes. They are also truly intermediate—"in the middle"—between coal-tar crudes and a host of drugs (as aspirin), plastics (as bakelite), chemicals for curing rubber, toughening varnish, thickening lubricating oil, purifying gasoline, killing insects.

The Dyes. How are dyes made? We cannot tell thousands of stories as one! Probably 5000 different synthetic colors have been made in laboratories; at least 1500 of these proved suitable for the coloring of textile fabrics. Each is made in its own way, although many are close kin and are the same up to the final processes.

In general a dye factory contains large tanks, partly filled with liquid or dissolved intermediates. To these, acids or alkalies, or other intermediates, are added under carefully controlled temperatures and stirred until they unite. Or dry intermediates and other dry chemicals are heated in kettles until they join in a molten mass.

The impure dyes formed by these reactions are soaked out by water, alcohol, or other solvent, then turned into crystallizable products by adding acids, alkalies, or salts of metals such as tin or zinc. The actual crystals are obtained by careful evaporation and drying. Dyes have various crystal forms, although most of them are ground rather fine, so they may dissolve quickly when used. It is interesting to note that the colors of powdered dyes are often quite different from their colors in solution.

Dyes are shipped in water-proofed packages, varying in size from one-ounce bottles to hundred-pound kegs.

CHEMICAL STRUCTURE OF DYES

Just as the physicists finally solved the mysteries of the rainbow's colors as wave lengths of light, so the chemists have learned the color secrets of the synthetic colors as due to arrangements of atoms within the dye's molecules. The color-heart of a dye is called a *chromophore*, Greek for "color bearer." There are but a few of these—about 15 in all. Every chromophore contains at least one double bond, as between two nitrogens of the azo dyes, between nitrogen and oxygen in the nitro dyes, between carbons in the colored quinones. There is some indication that these double bonds, being electrical in nature (have you studied about valence electrons?) permit electrons to vibrate from one atom to the other, "in tune" with waves of light, adding to the energy of some and subtracting from the energy of others, thus reflecting or absorbing color from the light.

The chromophore heart of a dye must have a body; it is the *chromogen*, Greek for "color maker." This is the simplest colored compound that can actually be made from intermediates. In this combination one or more benzene rings are attached to the chromophore, and each arrangement forms a definite chemical family of dyes.

Then on each body of a dye the legs and arms are fastened; these are chemical groups known as *auxochromes*, Greek for "color helpers." There may be one or several of these additions. They cause the differences between the color members of dye families, distinguishing them as the legs and arms, ears and noses, hair and chins help us to tell human brothers and sisters apart. There are many auxochrome groups, such as $-NH_2$, $-OH$, $-OSO_2H$, $-COOH$, Br, I; their presence brings out the true color of the dyes.

Dyes of simple structure are yellow. As groups of atoms are added as auxochromes the color deepens, and turns an orange hue because of developing red. As the molecule becomes more complex it becomes in turn a red, a purple, a blue, a green. Any dye molecule, wherever it is in this lineup of colors, is shifted along the line toward the green by the addition of even one more auxochrome group.

THE NAMES OF DYES

Every dye is a definite chemical substance, and is named precisely according to its structure by the rules organic chemists follow. These names are likely to be long—sodium carboxydimethylaminoazobenzene is a specimen.

Each dye also has a "trade name," which is sure to be much shorter; methyl red is the dye just mentioned so jawbreakingly. Unfortunately there is no precision in the trade names of dyes. Some are named after people—Bismark brown; or places—Congo red; or one of the chemical groups—methylene blue; or an important ingredient—sulfur black; or the color—nigrosine; or an important use—butter yellow; or a resemblance, as to an ore of copper—malachite green. The older dyes of German invention are likely to have names followed by a letter indicating a color or at least a shading of color, as Bordeaux R (*rot*, German for red); orange G (*gelb*, German for yellow); rhodamine B (*blau*, German for blue); and so on. The letters may be mingled; BG stands for blue-green. The letters may be repeated; benzopurpurine BBBB is of blue, bluer, yet bluer, and bluest shade! Still other letters are merely factory designations. A single dyestuff may be on the market under a dozen different names. Although most of these names sound pretty when pronounced properly, they are confusing and give little worthwhile information.

Still a third name—the fashion or popular name—may be applied to a dye after it is on the cloth. These names vary with the years and the seasons and the moods of purchasers. Such terms as sunset, straw, navy blue, grasshopper, midnight, red fox, geranium, paprika, stardew, buttercup, rhapsody may be crudely descriptive—but your guess as to the real color is as good as that of anybody else. It is really a pity that such an unscientific method of naming dyes and their colors has become a fixed custom. How much better it would be to call the tint of a brilliant scarf, not by its fashion name *sunset*, nor its trade name *Biebrich scarlet*, but by its chemical, and only correct, name—simply *tolylazotolylazobetanaphthol!* Do *you* agree?

HOW "NEW DYES" ARE INVENTED

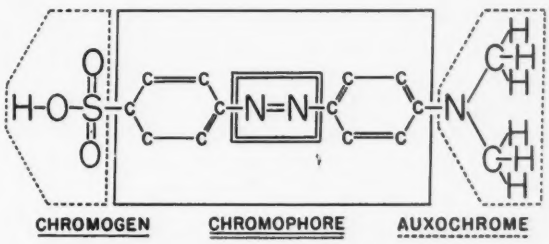
The dye chemist of today understands the inside construction of a dye molecule. He knows what chromophores he can begin with, and the actual colored compounds or chromogens that have been made already from each. He can experiment to make new chromogens. To these—old and new—he can attach auxochromes here and there to their structures. When a new dye is thus "invented" it is tested on the different kinds of cloth to determine its color strength and its fastness. A dye is called "fast" if its color *does not* run when washed, fade in sunlight, or change color from sweat, dust, or hot irons. A dye unsatisfactory in these respects is "fugitive"—it *does* run in water, fade in sun, and so on.

If a dye is proved good by tests, the cost of making fairly large quantities of it is estimated; if this proves economical, the new dye is put on the market, samples are sent to the dyers, and advertisements are inserted in the trade papers.

DYES CLASSIFIED AS TO USE

Practical dyers think little of the structure of a dye, but much of its effect on cloth. Their classification of dyes as to use is the one listed in Government reports, such as that issued annually by the U. S. Tariff Commission. Using the figures for 1939 in order of amounts, the following kinds of dyes were produced in United States dye factories:

Direct dyes, 31,438,000 pounds. These dye cotton directly, hence the name. They are also called *substantive dyes*, the word referring to a *noun* which names its object directly. Even dyers know grammar! Direct dyes are at the head of the list in pounds because of the vast amount of colored cotton worn in our nation.



CHEMICAL FORMULA OF A DYE (METHYL ORANGE)
Showing the *chromophore* group, enclosed in double-line rectangle; the *chromogen* group, in single line rectangle; and *auxochrome* groups, in dotted lines.

Vat dyes, 30,035,000 pounds. Only slightly behind the direct dyes in quantity manufactured are these exceedingly fast vat dyes. They are colorless when cotton or rayon cloth is first dipped into them, but are

oxidized in large vats by chemical oxidizing agents. The dye is actually made inside the cloth threads by this two-dip process (hence the common phrase "dyed in the wool," implying fastness). Indigo (from coal tar, not plants) for overalls is the chief vat dye. Its color comes by oxidation when merely hung in air.

Sulfur dyes, 18,651,000 pounds. These contain sulfur, hence their name. They dissolve only in cold sulfuric solutions, and settle out inside the threads when boiled, giving to cotton cloth an inexpensive dye exceedingly fast to washing.

Acid dyes, 17,700,000 pounds. These are the favored bright dyes for wool and silk. These fibers are proteins, and have basic groups related to ammonia in their structure that seem to have special attraction for acid dyes. (See the later discussion of the Theory of Dyeing.) Acid dyes thus dye wool directly, but can be made to unite with cotton by the use of mordants (Latin word for "biters"), in which "the mordant bites the cotton with its back teeth and

Courtesy of E. I. du Pont de Nemours & Company

STANDARDIZING LABORATORY

A yellow dye is being applied to a cotton skein, and the results checked to make sure that the dyestuff conforms to the rigid requirements of its standard in all particulars, before shipment is made to a customer.

the dye with its front teeth, and holds fast!" The mordants used with acid dyes are bases, usually jelly-like hydroxides of aluminum, antimony, chromium, iron, tin, which soak into the fibers. Because these jellies are bases, they unite readily with acid dyes.

Acid dyes are particularly fast to light, hence are used for outdoor decorations. Their fastness to washing is not the best. For buntings, flags, and similar decorations for parades and festivals, the bright, unfading acid dyes are often put directly on cotton and dried without washing. If it rains, what matter if the dyes *do* run—the celebration is a mess anyhow!

Basic dyes, 6,415,000 pounds. These are the deep, often dull dyes for wool and silk. The proteins of these fibers have acid groups related to acetic acid in their structure, hence basic dyes unite with wool and silk directly. Basic dyes also unite fairly well with cellulose materials such as cotton cloth and paper. They dye cellulose better if an acid mordant is used, such as tannic acid and tartar emetic (potassium antimony tartrate).

Basic dyes on wool or silk are usually very strong, giving more color per ounce than any of the other classes. They have a weakness, however—poor fastness to light.

Mordant and chrome dyes, 5,237,000 pounds. These dyes are neither acid nor basic enough to unite with the respective basic and acid groups of proteins, hence they must be attached by mordants—"the biters"—to wool and silk. The most effective mordant is potassium bichromate, called "chrome" by dyers. Chrome may have many "assistants," such as cream of tartar (potassium acid tartrate), oxalic acid, formic acid, lactic acid, and others. Less used with these dyes are the metallic mordants, jellies of iron and aluminum hydroxides.

The mordant dyes on wool are very fast, and for the deeper colors, particularly the blacks, these dyes are the best to be had.

Azoic dyes, 3,318,000 pounds. These interesting colors are actually produced in the fibers themselves, hence are often called *developed dyes*. Because the dyeing is carried out in ice water they are sometimes called *ice colors*. All of these dyes contain the double-nitrogen group as chromophore, hence their usual class name is *azo dyes*, *azoic dyes*, or *diazo dyes* from the French word *azote*, "nitrogen."

Azoic dyes are particularly useful for cotton. The cloth is first dipped in an intermediate containing one nitrogen. This is *diazotized* (its nitrogen doubled) by nitrous acid from sodium nitrite in a bath kept cool by floating ice. Then, in a third bath, another intermediate is coupled to the diazotized molecules, and a very fine color develops instantly. A great variety of colors is available.

Lake and spirit-soluble dyes, 3,305,000 pounds. If dyes are combined with mordants, not in cloth but in vessels, *lakes* are formed. These are insoluble in water, but will spread through oils to color paints and inks, while some will dissolve in alcohol (spirit) to color varnishes and lacquers.

Acetate dyes, 2,585,000 pounds. The cellulose acetate fibers called celanese, or acetate, are poorly colored by the dyes that work well on cotton and viscose rayon. New dyes had to be invented for these synthetic (man-made) fibers soon after they were invented by chemists. These dyes unite directly with the acetate group in the fiber, and will not give good colors on either cotton or viscose rayon.

Other classes of dyes, 4,824,000 pounds. A con-

siderable variety of dyes are made in small quantities each year, each having some special value for some specific purpose.

THEORY OF DYEING

How do dyes become fastened in the fibers so they will not wash out? The answer probably is—dyes are *colloids*.

When dissolved in water a dye appears to be a clear solution, with its particles so small that not even a microscope will reveal them. If a strong beam of light is passed sideways through the colored fluid, however, the cloudiness called "the Tyndall effect" may be seen. This effect is due to the scattering of light by particles that are of *colloidal* size—smaller than the particles of a suspension visible to the eye (or even a good microscope), but larger than the molecules of a true solution. The word *colloid* is Latin for "like glue," and you are familiar with colloids not only in glue but in egg white, clays, milk, jellies, and countless other familiar things, including soap (we hope!).

Now if dye particles were as small as those in true solutions, they would wash out readily. If they were as large as even the finest dusts they would dust out of the cloth when dry. But colloidal particles are *adsorbed* on many surfaces—spread out in thin layers that stick very tightly. The cause of this sticking is probably electrical, for surfaces and particles are covered with electric charges.

The word *ad*sorbed, which means "stuck against" in Latin, must not be confused with the word *ab*sorbed, which means "suck away."

Here is the picture the dye experts have of dyeing. When a fiber is colored directly (as direct dyes on cotton and rayon, acid and basic dyes on wool and silk) the fiber adsorbs the dye particles. These tiny fibers, swollen in water, open up many microscopic canals into which the particles—if not too large—will wander, and —if not too small—will stick in the thin layers that characterize adsorption. The staining of plant and animal tissues with dyes so that some parts will be colored and others clear under the microscope is an example of selective adsorption. The fact that many dye particles are washed off these surfaces if the cloth is soaked in alcohol is evidence that there is no chemical union with the fibers.

Mordanting must be done if the surface of a certain fiber will not adsorb the particles of a certain dye. (See previous paragraphs on mordanting.) The mordants are themselves colloids, jellies of metallic hydroxides for use with acid dyes, and glue-like tannates, etc., for basic dyes. The mordants adsorb on the fiber surfaces, and in turn adsorb the dye particles. One might compare mordanting with an old practice of sticking bright pieces of colored glass to a stone wall through the use of a thin coat of plaster. In this example, which represents the fiber—the wall, plaster, or glass? Which the dye? Which the mordant?

Colloidal particles run together and form larger clumps when mingled with solutions of electrolytes—salts which have charged ions in solution. This clumping, which is called flocculation, takes place rather slowly. It is a practice to add common salt to the solutions of direct cotton dyes, and certain others, to flocculate the dye particles after they have entered the canals of the fibers, making them too large to float out again.

Although chemical action of dyes with fibers is not ruled out entirely, most dyeing is well explained by the theory of adsorption by the fibers of dye particles of colloidal sizes. There are chemical actions of significance in the dye baths, however. The best examples are in the dyeing of wool and silk with acid and basic dyes.

The acid dyes are really salts; we may write a type formula as NaDye. For use, this salt in solution is "broken" with acid; a type equation is:

$$2\text{NaDye} + \text{H}_2\text{SO}_4 \rightarrow 2\text{HDye} + \text{Na}_2\text{SO}_4$$

The acid dye becomes the dye acid, and particles of this dye acid enter the wool and silk fibers. If deeper dyeing is wanted, sulfuric acid is used; for lighter shades the weaker acetic acid is employed. To make the dyeing slower so the color may enter the cloth evenly, sodium sulfate (called "glaubersalt" by dyers) is added. Since it is on the right side of the equation for the reversible reaction, the reason for its retarding action is obvious.

The basic dyes are also salts; a type formula, DyeCl. These dye salts are hydrolyzed by the water in which they are dissolved; a type equation:

$$\text{DyeCl} + \text{HOH} \rightarrow \text{DyeOH} + \text{HCl}$$

The basic dye becomes the dye base, and this enters the wool and silk fibers. For even dyeing, hydrochloric acid—seen on the right side of the equation—is added to retard the action.

DYEHOUSES

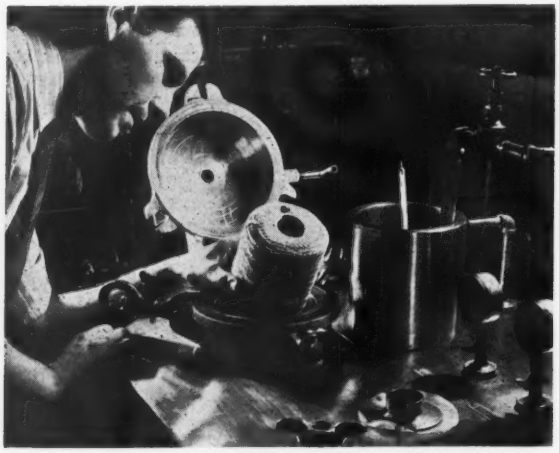
Textile fibers are dyed either *in the yarn* as threads, or *in the piece* as woven cloth. These receive their color in buildings called dyehouses. What is a dyehouse like?

Contrary to what one might expect, a modern dyehouse is not a sloppy place with colors splashing over floors and walls. Care and cleanliness are necessary for level (even) dyeing, and precise quantities of everything used in the dye bath are measured out. Waste is expensive, and must be avoided. The costs of dyestuffs themselves vary from 50 cents to $5 a pound.

The top floor of a dyehouse contains a well-lighted room with large windows, or a skylight, so that natural light is available most of the day. It should be well ventilated to carry off vapors from warm baths. Modern yarn dyeing machines are tanks of stainless steel, with stirrers, into which the dye liquor is pumped at the right temperature. Elsewhere in the room is an extractor to which the yarn, wet with dye, is carried and whirled rapidly. Much of the liquid is spun to the outside of the extractor by centrifugal force, and drains

away. Where it is necessary to wash the dyed yarn, the extractor removes the wash water. The rest of the moisture is taken off in chambers of hot air.

Piece dyeing machines usually carry the cloth over rollers into and out of a dye vat. The cloth may be washed in similar vats, the water squeezed out between rollers, and the drying done in warm air chambers.



Courtesy of E. I. du Pont de Nemours & Company

PACKAGE DYEING MACHINE

A considerable amount of yarn is dyed in what is known as "packages." These are steel coil bobbins wound with yarn which are placed over circular tubes in a cylindrical machine, the cover of which is securely bolted. The dye liquor is pumped through these perforated tubes and forced out through the cones. The machines in the illustration are small size for laboratory use, but are in every essential respect like full size models.

In the basement of a dyehouse is a water softener—for minerals in a dye bath may greatly alter the colors. A water heater supplies the necessary warmth of most dye baths. Cloth is usually entered at moderate temperatures, and gradually heated up to boiling. A brine tank supplies salty water when needed, particularly for the direct dyes on cotton. A tank of bleaching chemicals permits a color that is stronger than desired to be made lighter.

At a convenient location in the dyehouse is the drug room where dyes and other chemicals are weighed out. There is also a laboratory where new dye mixtures for new colors are tried out in a small dye machine, and an office where records are kept.

How much dye is put in a dye bath? No cloth in a dye bath will take up *all* of the color. A principle of physics called the *law of partition* seems to apply; for example, if cloth takes one-tenth of a dye in a bath the first time, the same amount of cloth would take one-tenth of what is left the next time, the third quantity of fresh cloth would take one-tenth of the remaining color, and so on. It is also true that cloth takes a definite quantity of dye the first minute it is in the dye bath, less the second minute, still less the third minute, and so on. The degree to which color can be used up is called the *exhaustion* of the dye. Some dyes exhaust better than others; animal fibers (wool and silk) exhaust dyes better than vegetable fibers (cotton, linen, rayon).

In the modern dyehouse the concentration of baths is kept at a proper level by adding new dye solution continuously through the pumps. The amount of dye put in the baths is calculated on the weight of the cloth to be dyed. For light colors $^1/_4$ to $^1/_2$ per cent of dye is used—that is, $^1/_4$ to $^1/_2$ pound of dye for each 100 pounds of cloth. For strong dyeings 1 to 3 per cent of dye will be needed. Very deep dyeings as of blues, browns, and purples require 5 per cent of dye. Most of the black pieces of cloth are not dyed with black dyes at all, but with 7 to 10 per cent of green, blue, brown, or purple dyes. It would of course not be possible to get black cloth with yellow, red, or orange dyes.

SMALL-SCALE DYEING

Not all dyeing is done in large dyehouses. Many small dye establishments are busy, chiefly in giving new colors to garments already made and dyed. In this work the dyer first determines the nature of the cloth, for his procedures are quite different with the animal, the vegetable, and the synthetic fibers. Then he places the garment in a vat, soaks it in warm water and may wash it in soapy water. After rinsing he may bleach the garment in a kettle filled with bleaching solutions, heated to the right temperature with steam coils.

Then he fills a copper dye kettle with hot water, carefully measures his dye on a basis of the weight of the goods, and stirs the color in. He may use one dye or a blend of several dyes. He turns on steam to heat the solution, drops in the garment, and stirs it with a wooden stick. From time to time he pulls the garment out, dries one small spot with an iron or hot air hose, and compares the spot with a sample of the color he desires. The drying is necessary because wet cloth appears darker than dry cloth. If he needs more color, he may let the garment stay longer, or he may add more dye. If the color is too deep, he has bleaching solution ready to pour into the kettle.

Satisfied with the color, the dyer lifts the garment out and "fixes" its dye either in a cold-water rinse, or fixing chemicals. These are soap, salt, or other chemicals, depending on the fiber and the dye. The garment is then partially dried in a centrifuge extractor, and completely dried in warm air.

These small dyeing operations are usually associated with cleaning and pressing establishments.

HOME DYEING

"To dye or not to dye—that's the question!" Home dyeing is usually a renovation of fabrics or garments which may be of further service. Faded colors of underwear or summer clothing may be renewed by bright colors. A mother's green dress may not suit the daughter, but when dyed navy blue it may be

made over into an attractive "new" frock. One may merely become tired of curtains, cushions, and table covers which are easily given a fresh color scheme by dyeing.

Home dyeing is conveniently done with packaged dyes, bought in envelopes, tubes, boxes, or soap cakes at almost any drug and department store. These packaged dyes are usually blends of two or more dyestuffs, hence are of great variety in color. They are available in the changing popular shades of the seasons. They are often toned by small additions of red dye to "warm" the shade, or blue dye to "cool" it, by black dye to add gray tone, yellow to brighten, etc. This may be proved easily by sprinkling very small quantities of the packaged dye powder on the surface of water, and observing the spots of different colors. A very few packaged dyes include but one dyestuff.

One package of dye colors one pound of material deeply, and four or five pounds lightly. Full directions for use are given in each package, and it is not necessary to repeat them here.

Several things must be considered when selecting a packaged dye for use at home:

1. *The kind of fiber to be dyed.* Certain packaged dyes are for wool and silk only, others for cotton and rayon only. Certain packaged dyes are for both animal and vegetable fibers; these are called *union dyes*. Their effect on synthetic fibers, the rayons, should be tested before use.

2. *The shade desired.* One rule of home dyeing is this—always dye to a darker shade. Light colors may be covered easily, but strong colors must be dyed much stronger, or black.

3. *The previous color will affect the new color.* Purposely blending the old and the new is called *top dyeing* —that is, dyeing on top of other colors. Blue over yellow makes green, green over red makes brown, etc.

If the old color is too strong, and not wanted, it may be bleached by boiling with chemicals known as *dischargers* or *strippers*. This must be done most carefully with wool and silk, and the chemicals are likely to weaken cotton.

One must be prepared to find some shrinkage of redyed materials. For this reason well-fitted men's suits are rarely dyed. Ladies' garments are frequently ripped entirely apart, and after dyeing resewed to new measurements. This assures proper fit and new, strong seams.

Home dyeing has many artistic variations, and becomes an interesting hobby to many. Tie dyeing and batik dyeing are among the most interesting applications.

Tie dyeing, a very ancient craft of Asia, is done by winding and tying parts of the fabric into tight lumps that will not be dyed. The color enters the looser portions. The cloth may then be partly unwound for another, and darker, color—then still further untied for yet darker colors. When the knots are finally loosened and the cloth spread out, interesting symmetrical patterns are seen.

Batik dyeing originated in ancient Java. Silk is preferred for this craft. Wax (beeswax and paraffin) is melted, and brushed or blown on the cloth, covering the parts—from broad areas to fine lines—that are not to be dyed. Then the lightest color is applied in a cold bath, after which the silk is rinsed and dried. The wax is removed by ironing between blotting papers or dipping in gasoline. New areas—colored or uncolored—are waxed to resist the second dyeing in the next deeper color. A repetition of this process, with much patient skill, gives designs of pleasing colors and patterns.

"UNDYEING" CLOTH

Taking color out of dyed cloth is called *discharging*, or *stripping*. Even good work must sometimes be undone. The simplest need for "undyeing" arises least often—that too deep a color has been applied in the first dyeing of cloth. Only a careless dyer lets this happen. Practically all coal-tar colors may be removed in warm ammonia solution. The dyer who has overshot his color shade puts the cloth in ammonia until the color is considerably toned down, then dyes again—and resolves to be more careful.

Courtesy of E. I. du Pont de Nemours & Company

LAKE LABORATORY PICTURE

Dyes are not only used for the coloring of textiles, paper and leather but are also employed to a considerable extent in imparting color to printing inks, paints, varnishes, lacquers, wallpapers and so forth. Depicted is laboratory equipment used in the conversion of water-soluble dyes into insoluble pigments, known as "lakes."

The next need for discharging color comes also from error, though not from carelessness. A manufacturer may find himself with a large stock of goods in an unpopular shade, due to the fickleness of fashion. He

ships this material to a dyehouse to be stripped of its color and redyed to an acceptable hue of the season.

The strippers used to discharge colors are strong reducing agents. Some are relatives of sodium hydrosulfite, $Na_2S_2O_4$; one of the most powerful is sulfoxalate, $NaHSO_2 \cdot CH_2O$, a combination with formic acid. Other strippers are organic derivatives of ammonium salts in which all four hydrogens of NH_4Cl are replaced by methyl(—CH_3) or other organic radicals. Most of the dischargers or strippers are patented mixtures sold under trade names.

The third and chief use of strippers is in *discharge printing* of cotton cloth. If a printed design has more white than colored area (such as blue dots on white cloth) the dye is printed on by rollers in the manner of a printing press. If the design has less white than color (as white dots on blue cloth) discharging chemicals are printed on cloth of solid color. These powerful chemicals sometimes weaken the cloth, and holes come in the white lines or dots.

Resist dyeing is also carried out to form white designs on colored cloth. The principle is much the same as batik dyeing. A wax is printed on the cloth which resists the dyes applied in the dye baths. After washing the wax is removed by heat or gasoline. Fine white dots and lines are easily put into a design by this plan.

THE FASTNESS OF DYES

Manufacturers of dyestuffs make many tests of their products, and tell the dyers in all honesty exactly how well the colors should endure when the cloth is used. Colors that do not fade are called *fast*, those that fade are *fugitive*.

Fastness to light was formerly determined by weeks of exposure to sunlight, but is now tested in a few hours under the rays of an ultra-violet lamp. The manufacturer states the quality of his dye as to fastness to light in the terms excellent, good, moderate, fair, and slight. Dyes fast to light should be selected for curtains, awnings, outer clothing, and similar materials that will often face the sun.

Fastness to washing is tested by putting white and colored cloth, braided together in strips, in hot, soapy water. If color is transferred to the white piece the dye is not fast. The manufacturer gives the information for each dye as to whether white goods are untinted by heavy, medium, or light dyeings, or slightly tinted by heavy, medium, or light dyeings. The best of these dyes should be selected for underwear, table furnishings, towels, and other cloth that must go often to the laundry.

Fastness to acids is tested by soaking in a weak acid—acetic—and a strong acid—hydrochloric. After washing and drying the test sample, its color is compared with the original samples, and any change in the shade is noted. The dyes most fast to acid are unaltered in hydrochloric acid; dyes less fast are reddened, blued, blackened, etc. Dyes of moderate fastness to acids may be unaltered in acetic acid; poor dyes are reddened, blued, blackened, etc., in this weak acid. Outer clothing worn in the atmospheres of smoky cities, or which are subject to perspiration, should be colored with dyes of much fastness to acids.

Fastness to alkalies is tested by splashing the cloth with drops of a mixture of lime water ("milk of lime") and ammonia. After a short time the cloth is rinsed and dried, and any lack of fastness is shown by spots in which the color is reddened, darkened, greened, lightened, etc. If the dye is fast the shade is unaltered. Clothing that will collect street and road dust should be colored with dyes fast to alkalies.

Fastness to hot pressing is tested by running a very hot iron over the damp cloth. If the dye is not fast the color will change, and keep its new tint on cooling. A fast dye may change color when hot, but the original color will return when it cools. Manufacturers refer to their dyes as excellent, good, moderate, fair, and slight in their fastness to hot pressing. Since much cotton cloth is calendered, or pressed between hot rollers during its finishing process, and since the iron is both a laundry and a household necessity, dyes should be at least good in their fastness to hot pressing.

Fastness to hydrosulfite discharge is important to the dyer of goods that may have to be bleached and redyed, or finished in a color design with white areas obtained by discharge printing (see previous paragraphs). For each dye the manufacturer tells whether it is discharged with ease, with difficulty, or not at all, and whether the bleached areas will be truly white or tinted.

DYES AS INDICATORS

Every student of chemistry is familiar with *indicators* that are one color in acid solutions and another color in alkaline solutions. Some indicators are vegetable dyes (litmus is an example), but most of them are synthetic organic colors—coal-tar dyes. Acid dyes that are indicators become dye acids of one color in acid solutions, and dye salts of another color in basic solutions. Basic dyes become dye salts in acid solutions and dye bases in basic solutions.

Certain dyes (the pH indicators) change color rather quickly, and are said to have a "sharp end point." They are used for testing very weak solutions of acids and bases, such as found in plant juices, medicines, the body's fluids, and the like. Few indicators change color at exactly the neutral point between acid and base; for some dyes the mid-point is slightly acid, and for other slightly basic.

COLORS IN FOODS

Soon after the coal-tar dyes became numerous the bakers and candy-makers of Europe began to color their products with the new and powerful chemicals. Soon it was learned that many of the dyes were poisonous, or at least upsetting to the human digestion. Gradually "approved lists" of harmless dyes were made. After the passage of the Pure Food and Drug Law of 1906 in the United States, the use of any harmful dye in foods was prohibited. Tests were carried out,

first on dogs, then on "poison squads" of young college students. At the conclusion of the tests, only seven colors out of the several hundred previously used in foods were proved harmless. These were listed as "Certified Food Colors," and their use permitted by law unless used to conceal poor quality of the food. On packaged goods, such as soft drinks, catsup, mustard, and many others, the presence of the color must be mentioned on the label.

The original seven colors have become 16 by the addition of new harmless dyes. There are four reds, one orange, six yellows, three greens, two blues. Three of the yellows are soluble in oil, all others are soluble in water.

The Pure Food and Drug Law of 1906 was revised and enlarged in 1938 as the Federal Food, Drug, and Cosmetic Act. In 1942 the U. S. Department of Agriculture was given the responsibility of determining the safety of dyes used in drugs and cosmetics, and as this is written about 130 of them are being offered by manufacturers who give testimony as to their harmlessness. Within several months a Certified List of Colors for Drugs, and another for Cosmetics, will be published. Certification for use in cosmetics, however does not permit use in foods or drugs.

Within the past several years we have often seen a person wearing—usually with some embarrassment—a brilliant spot of red, blue, or purple on face or hands. The color is not for style, but for safety; the germ-killing power of certain dyes was discovered many years ago. If the spot is red, it is probably mercurochrome, which contains three benzene rings, two sodiums, two bromines, and one atom of mercury, together with a quota of carbons, hydrogens, and oxygens in its complex molecule. If the spot is blue-green, the protecting dye is probably methylene blue; if purple, it is gentian violet. These dyes, though milder than iodine in their germicidal power, are not irritating to the skin. The amazing family of sulfa drugs have been developed since 1932 from a red dye, prontosil, that had been found of unusual effect in killing germs.

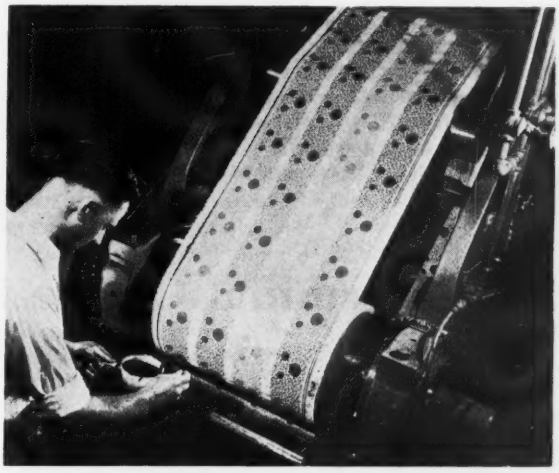
FLUORESCENT DYES

Certain dyes, chiefly of the fluorescein family, shine with amazing brilliance under the invisible rays of so-called "black light." This is short-wave ultraviolet energy, invisible to the eye. Under ordinary light these dyes show their colors of red, yellow, green, blue, violet, etc. In the dark they are as colorless as any other dyes. But the moment the "black light" is turned on their fluorescent colors startle all observers with their brilliancy!

These dyes have been used much for stage effects, for trick pictures that change with the flip of an electric switch, and for the "glowing carpets" of motion picture theaters.

MISCELLANEOUS USES OF DYES

Almost every discussion of a topic that has many sides comes finally to a heading of "Miscellaneous." This means that the end is near. It might mean that the most important matters have been presented, and here are some crumbs to be swept up. In this the miscellaneous paragraphs may give the wrong impression, for a "crumb" of comment may concern very important information. Frankly, "miscellaneous" is an author's challenge to a reader to seek further knowledge; "my allotted space is about filled and I have given your interest a good start. If you want to know more, here are topics you should study, and if you seek this information you can find it!"



Courtesy of E. I. du Pont de Nemours & Company

PRINTING MACHINE

Testing dyes for textiles printing. This machine duplicates exactly, for experimental purposes, the operation of textile printing and clearly demonstrates the suitability of dyes for this specialized form of application

Dyes in *writing inks* are usually dissolved in water. Red ink is usually eosin, green is malachite green, blue is methyl blue, black is nigrosine or acid black. *Marking inks* are usually aniline black, which develops color on exposure to air. *Copying inks* and the dye in the center of a copying (or indelible) pencil are usually methyl violet for purple writing, victoria blue and malachite green for two other colors.

Typewriter ribbons are colored with methyl violet (purple), nigrosine (black), eosin (red), victoria blue, or malachite green. These colors are mingled with glycerine before spreading on the ribbon. *Carbon papers* are coated with the same dyes, held to the paper with fine gum.

Colored *printing inks* mingle their colors in linseed and other oils. Mineral pigments are much used (see discussion of the mineral colors). When synthetic organic colors are employed they are turned into *lakes* by union with metallic hydroxides. Both pigments and lakes, being insoluble in oils, are ground to extreme fineness—even to the size of colloidal particles. Black printing ink is not made from a dye, but from carbon black made from the smoky flames of natural gas.

Paints are also colored with mineral pigments and dye lakes, which color the particles of white lead and white zinc that give the covering power to paint. In *varnishes* dyes color the colloidal gums that are dispersed through the varnish fluid—usually methyl alcohol; the dyes selected, therefore, must be soluble in spirits (alcohol).

Soap is colored by selected dyes that do not change color with weak alkalies. The color is added while the soap is still melted. In a similar manner the colors of *candles* are mixed with the melted wax. Dyes are also specially prepared for solution in *gasoline* to distinguish the different grades of motor fuel in the pumps.

Much colored *paper* is used; if the sheet is colored on both sides alike, the dye has been mingled in the pulp. Paper colored on one side has been printed with ink.

Dyes for modern *synthetic plastics* are added to the molding powder. They should be particularly fast to light. The same fastness is necessary for the dyes blended in the *synthetic lacquers* that gaily color our automobiles.

Color photography has been made possible by the invention of methods by which dyes are substituted for silver particles in exposed and developed sensitive emulsions on the backs of celluloid films. Three colors—a red, a blue-green, a purple—blend to furnish all the close reproductions of natural colors in the finished color films, such as kodachrome and dufaycolor. The still pictures that the camera club members show on a screen from small slides, as well as the amazing technicolor that flashes over the screens of our movie theaters, depend alike on tiny areas of dyed film.

Any other uses for dyes? Let the writer think—hard! Yes—here is an unmentioned service—wax-soluble dyes in *shoe-polish*. This brings us to the toe-tip of our story of synthetic organic colors, and seems a good place to stop.

A Terminal Course in General Chemistry[1]

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A TERMINAL course in chemistry is a course which leads to some comprehension of the whole field of chemistry. It is a course which teaches the methods, the achievements, and the problems of chemistry without planning to turn out trained professional chemists. The course capitalizes on the great range and fascination of chemistry at the deliberate expense of neglecting preparation for more advanced work. It is assumed that few of the students will take a second year course in the subject.

The need for such a course has been apparent for a long time. Some of our better elementary chemistry textbooks have gone so far as to introduce a chapter or two on the chemistry of carbon. A few survey type books have been written, but the writer knows of no case in which such a course has been an outstanding success. Many such courses have been abandoned after a few years' trial. Nevertheless the need for such a course remains. Few students elect a year of general chemistry for the intellectual satisfaction that it offers. Those who do often complain about the technical and dull details that go into the training of a professional chemist. The necessity for a terminal course therefore exists, but the problem is not solved by admitting the need. The problem can be approached only in terms of student preparation and the relationship of the course to the curriculum as a whole.

The students who register for elementary chemistry differ widely in preparation. Some have had good high-school chemistry courses, some poor, some none. The students differ widely in interest; some are intensely interested in chemistry, some bored to tears. Some students plan to be professional chemists, but the great majority do not. Some students are capable of becoming scientists, many are not. There exists, therefore, a great difficulty of classification, and immediate segregation of the supposed scientists from the others is likely to lead to dissatisfaction, failure, and to a general lowering of morale.

It is suggested, therefore, that all students be put into a first quarter of general chemical principles, but that they be segregated into sections for those who have and those who have not had high-school chemistry. In the writer's experience the latter group needs an extra lecture per week, chiefly to gain familiarity with the principles that the laboratory work is designed to illustrate. This first quarter is a fairly rigorous introduction to the atomic theory, the kinetic-molecular hypothesis, and the periodic table. The course is difficult, and replete with arithmetical problems, and purposely so. At the close of the first quarter, the larger group of students, which plans more than one year of chemistry, and has proved itself qualified, is invited to register for the remaining two quarters in general chemistry and qualitative analysis, which are treated from a strictly professional standpoint. The remaining students go into the terminal course, which is our present concern. The first quarter serves therefore not only as an introduction to chemical principles, but as a comprehensive and reliable placement test. This plan is now in operation at Northwestern and it gives evidence of being highly successful.

[1] Presented before the Chemical Education Group, Chicago Section of the American Chemical Society, March 20, 1942.

The students who elect the terminal course are not, as might have been anticipated, the poorer students, although the average grade is slightly lower than that in the professional group. The course generally attracts a few of the very best students in elementary chemistry. These are students who are majoring in commerce, or political science, or some other non-science division. The course attracts many of the women in the department, but even so, women students have never made up as much as 50 per cent of the class. The course also contains a few individuals who had originally intended to be chemistry majors, but who, because of their experiences and grades in the first quarter's work have thought better of it.

There is every expectation that as the course becomes better known more and more good students will wish to elect it. For instance, this year, in the face of reduced enrollments and emphasis on practical training, the terminal course has slightly more than held its own in relation to the professional course.

There are certain pitfalls to be avoided in the organization and advertising of the terminal course. One of these is the impression that the course is designed for the poorer students. Before this pitfall was recognized, in an earlier year, the course was so predominantly made up of C and D students as to create a serious drag on teaching efficiency, if not an academic hazard to the instructor who had to report so many low grades. Now, however, means are taken to combat this difficulty and these have proved successful. It is necessary for all instructional staff in contact with elementary students to emphasize that the terminal course and the professional course are parallel in everything but content. It is necessary to make all tests and examinations at least hard enough so that the label "snap-course" cannot be applied. And finally, it is necessary to make the course so interesting that even the best student will find ample scope for his abilities.

A course in chemistry may have many objectives. It may, for instance propose to push back the students' horizon regarding the "nature of the physical world." It may try to inculcate the scientific method. It may prepare for more advanced courses. Pushing back the students' horizon is one of the valid objectives of the terminal course. The objective is achieved by suitable choice of subject matter. For a student who plans to take no more chemistry, the choice of subject matter should be as broad as possible.

Inculcation of the scientific method is also an objective of the course, but a minor one. It is necessary not to confuse the objectives of the course. Because an item teaches the scientific method it is not necessarily dull or lacking in interest to the lay mind. And because an item (like "heavy water") is fascinating to the lay mind does not mean that it cannot be used as a vehicle to drive home the scientific method. But after all, those who use the scientific method habitually have come to do so through many years of training. We can try to adapt our presentation and our laboratory work toward the scientific method. We should not feel too disappointed if the students show little progress in this direction. Clear thinking, to be sure, is the greatest asset a chemist can have. But scientists do not have a monopoly on clear thinking.

But no matter how many, and how admirable, the minor objectives of the course, the major objective of a terminal course in general chemistry should be to explain the impact of chemistry on modern civilization. There is no member of society who cannot use a rudimentary knowledge of chemistry. There is no thinking member of society who can afford to be ignorant of the effect of a Pasteur, a Haber, or a Hall, for good or evil, on the whole course of human history.

The course, as given by the writer, is still in the experimental stage. It consists of two lectures, one quiz, and three hours of laboratory work per week. The primary function of the lecture is to inspire the student to want to learn. Some instruction probably sinks in during the lecture, but the greater part of the actual learning process is expected to take place in the laboratory and at individual study. If the primary function of the lecture is inspiration, then the lecture must be very liberally illustrated with the finest kind of demonstrations. Use is also made of moving pictures and slides, and of microprojection. There are several sources of good demonstrations for physical and inorganic chemistry, but few for organic and biochemistry. For the major industrial operations, for nutrition, and vitamin and hormone deficiency diseases, liberal use is made of slides, colored, wherever possible. One rather critical test of how far the lecturer is succeeding in inspiring the students is to count the number of times students ask permission to bring their friends to class. Attendance at lecture is made optional, and no seats are assigned. Under these circumstances attendance is as good and class morale perhaps a little better than in the more conventional type of lecture. One reason for having only two lectures per week is that the staff has scarcely time to prepare more than twice a week for the kind of lecture desired, if the course is to be a success.

In the quiz period, questions are answered, discussion encouraged, and a short written test given every week. The purpose of the test is partly to keep the students up in their work, and partly to let the instructor keep his finger on the pulse of the class. In this way it is possible to make continuous adjustments of presentation and, by the method of sampling, to discover even in a large class, what things give the most difficulty and what should be altered or corrected for another year.

The laboratory is held at three hours per week. Longer laboratory hours take too much time from the non-science major and, if given, only succeed in teaching more manipulative technique, which is not one of the objectives of this course. The primary function of the laboratory is to lend reality to the subjects and materials discussed in lecture. The "scientific method" is also an objective which may be approached by skilful adaptation of the laboratory experiments performed.

But the approach to this objective remains in the exploratory stage so far as the terminal course is concerned. Student interest can also be aroused to a remarkable degree by appropriate choice of laboratory assignments.

The writer is familiar with one important Chemistry Department where the terminal course is offered without laboratory work. This is equivalent to fighting with one arm tied behind one's back. It arises, apparently, from the misconception that manipulative technique is the primary objective of laboratory work.

The next consideration is the choice of topics to be discussed in lecture. With the whole world of chemistry to choose from there must be compelling reasons for the inclusion of any particular topic. And, relieved of the necessity for preparing his students for a second year's work, the instructor may omit many topics commonly thought proper in elementary chemistry. For instance, it is neither necessary nor desirable that a student should be required to balance oxidation-reduction equations. On the other hand, it is highly desirable that he should know what will happen to him if he doesn't get enough vitamin D.

In order to weigh the importance of any given topic to the course, the following criteria have been set up:

1. Is the matter essential to understanding what comes next?
2. Can the student understand the matter with his background?
3. Is the student ever likely to come in contact with the matter again, either through his work, his reading, or conversation?
4. Does the matter concern the student vitally through his physical, social, or economic contacts?

It will be noticed that nothing is said about arousing the student's interest, and training his mind. Both of these worthy objectives are achieved by method of presentation rather than by course content.

The primary function of the laboratory work is to sustain the student's interest and to lend reality to the substances and reactions studied. It is not necessary that the laboratory work should closely parallel the lectures unless that can be done without sacrificing any of the other values in the laboratory work. The experiments should, of course, in a general way illustrate the lecture material. It is, however, important that the experiments should lead the student to a sense of accomplishment. To this end, experiments such as the analysis of an alloy and the oxidation of an alcohol seem to be highly desirable. Experiments such as the purification of sodium chloride, and the preparation of barium thioarsenate seem to be nearly useless. Another important feature is the use of problems. We might as well admit that many students cannot do arithmetical problems of any but the simplest kind. This is particularly true of students who elect the terminal course. The time necessary to teach them elementary arithmetic, not to mention algebra, is simply not available. It is desirable, therefore, to look for non-arithmetical problems, but problems which will, if possible, require the same kind of mental effort as do arithmetical problems. Fortunately there are such problems in organic chemistry. Especially valuable are problems requiring derivation of a structural formula through several steps of reasoning. Students enjoy problems of this type and are successful where they are admitted failures when faced with the arithmetical type of problem.

The tentative lecture syllabus which has been developed in accordance with these criteria is divided into two parts and covers the following topics: the non-metals, especially sulfur, chlorine, nitrogen, phosphorus, and silicon (oxygen and hydrogen are covered in the first quarter's work); colloids; the metallic elements, sodium, iron, magnesium, calcium, copper, silver, gold, zinc, mercury, aluminum, lead, tin, arsenic, with, of course, some reference to the less familiar metals; and important topics such as photochemistry and electrochemistry. This part takes about one quarter. The second quarter is devoted entirely to organic chemistry as follows: hydrocarbons, alcohols, ethers, aldehydes, ketones, acids, carbohydrates, amines, amino acids, phenols, quinones, heterocyclics, industrial aliphatics, coal tar products, and natural products.

Some idea of the syllabus will be gained by describing in more detail two of the topics discussed, one in the inorganic field, one in the organic. The first is the element silicon, which must be pretty well covered in one day. The lecture outline is as follows:

1. Relationships in Group IV in the Periodic Table
2. The Element Silicon:
   (*a*) abundance
   (*b*) central position in inanimate nature
   (*c*) preparation of the element
3. Compounds:
   (*a*) silicon tetrachloride
   (*b*) "Carborundum"
   (*c*) silica, quartz, "silicosis"
   (*d*) silicic acid
   (*e*) natural silicates
   (*f*) water glass
4. Glasses:
   (*a*) soft, Pyrex, Vicor, etc.
   (*b*) the ceramics industry

This, and all other lectures, are illustrated as copiously and as strikingly as possible. For instance, this lecture includes demonstrations of smoke screen formation, optical and thermal properties of quartz, formation of silicic acid, different types of glasses, and methods of glass-working. The more static type of demonstration should be, throughout the year, handled by means of semipermanent displays.

One of the lectures on Chemistry in Medicine covers the following topics:

1. Definition of Chemotherapy
2. Classification of Chemotherapeutic Agents:
   (*a*) analgesics, anesthetics, antipyretics, germicides.
3. General Anesthetics:
   (*a*) nitrous oxide, ether, chloroform, ethylene, etc.
4. Local Anesthetics:
   (*a*) cocaine
   (*b*) novocaine
5. Analgesics:
   (*a*) aspirin

(*b*) phenacetin
6. Hypnotics:
(*a*) bromides
(*b*) the barbiturates
7. Ephedrine and Benzedrine
8. Chaulmoogric Acid
9. Arsphenamine
10. Sulfanilamide and related drugs

The students are not required to memorize the formulas of complicated organic compounds, but many of them do simply because of their interest in the subject. One student even memorized the structural formulas of the important sex hormones. This proved a trifle embarrassing to the instructor who had to look the formulas up to see if the student were right. Demonstrations for this type of lecture are rather difficult and reliance must be placed on slides and motion pictures. However, the students get considerable satisfaction out of examining bottles filled with the various compounds.

It will be seen from this set of examples that the course is in no danger of being called a "snap course." In no sense is it designed for students who have proved themselves unable to assimilate chemistry. Rather it is designed for intelligent students whose major interest, for one reason or another, lies elsewhere.

The first half of the laboratory work is devoted principally to qualitative analysis. The reason for this is the markedly improved interest that students show when they are allowed to work toward a definite and realizable end such as analysis of an unknown, rather than being required to repeat miscellaneous inorganic preparations. With the limited time available the number of metals to be analyzed is restricted to lead, silver, copper, zinc, iron, nickel, calcium, and sodium, together with ammonium and a few representative non-metals. The procedure is adapted from the standard qualitative analysis procedures, with numerous modifications and simplifications. The metals chosen are those which are easy to detect, and which show a variety of reaction types. Semi-micro procedures are not used because the seven or eight periods of three hours each are too short a time to permit sacrifice of even a single period to filling reagent bottles and learning the semi-micro technique.

After his brief experience with qualitative analysis the student performs several experiments on photography such as developing and fixing a print; and on electrochemistry such as making a cell and doing some electroplating. Then he goes on to a series of experiments on organic chemistry. These experiments are:

1. Hydrocarbons, preparation, properties, identification of an unknown gas.
2. Fractional distillation of alcohol, oxidation products of alcohol.
3. Preparation of esters and ethers.
4. Preparation and properties of acetone
5. Preparation of aspirin.
6. Extraction of eugenol from cloves or caffeine from tea.
7. Isolation of a protein or a sugar from milk.
8. Reactions of proteins.
9. Miscellaneous reactions of Aromatic Compounds.
10. Synthetic polymers.

In the laboratory the students are allowed to work at their own speed, with occasional deadlines to round up the stragglers. Every effort is made to have the students think they are doing the experiments because they want to, rather than because they are required to.

Such a course presents obvious difficulties in regard to choice of texts. There are a very few texts which attempt to cover the field. For instance Hatcher "Introduction to Chemical Science" (Wiley), and Dyer "A Practical Survey of Chemistry" (Holt) both take a step in this direction. But neither of these texts, or any others which have come to the writer's attention are nearly advanced enough for this course. Furthermore they both have the sound of being written down to the level of students who are unable to understand the standard type of elementary chemistry text. So far, the writer has solved this problem by using two texts: Richardson and Scarlett "General College Chemistry" (Holt) for the inorganic part, and Fuson, Connor, Price, and Snyder "A Brief Course in Organic Chemistry" (Wiley) for the second part. This procedure is proving satisfactory except that the students have to buy two books. It is to be hoped that a text will be written combining the good features of both these excellent books.

The laboratory directions for the course are handed out in mimeographed form.

To date, student response to the course has been very gratifying although there remain some problems to be solved. For instance, there is the student who decides after all to major in chemistry. Our bulletin states that the terminal course does not admit to second year chemistry. In spite of this, an "A" student will generally be permitted to go on if he cares to do so. Other students who may wish to go on are required to take a course in qualitative analysis, for somewhat reduced credit. Another question is what to do with the very best students who will generally assimilate all that is given to them and ask for more. For the present an effort to solve this problem is made by putting all A and A− students in a special quiz section. Their instructor has directions to take them as far as they seem to want to go. For instance they will probably go into the theory of solubility products, while the rest of the class is content with a qualitative application of the law of mass action.

Faculty response to this course has also been very satisfactory. This is particularly true because it gives much more freedom to the professional course in elementary chemistry. One administrative officer suggested that the multiplication of freshman courses was likely to confuse the students. The answer is that the students don't seem to be half as confused as they were when they were all lumped together in one class.

The success of this terminal course seems to be reasonably assured. As various problems are worked out, and as the course becomes better known among the students, there is every hope of having it become one of the most popular science courses in the University.

What the Chemical Warfare Service Is Doing[1]

MAJOR WILLIAM O. BROOKS, C. W. S. *Washington, D. C.*

ALL AMERICAN citizens are interested in the activities of their Army and Navy, particularly at this time. Chemists, of course, are likely to have a special interest in the Chemical Warfare Service, because it is more closely related to their profession than are the other branches. As soon as our Service is mentioned, the average person thinks at once of chlorine and other poison gases, yet the program assigned to the Chemical Warfare Service is so inclusive and diversified that few civilians have any adequate conception of its ramifications.

The Chemical Warfare Service has a fourfold mission: research and development, production, instruction of chemical troops for combat.

This is a really big program and, even though poison gas has not yet been used to any great extent, the constant threat of its extensive employment places a heavy responsibility upon our Service, in order that we may not be caught napping. So many times we hear the question: "Why has poison gas not been used in this war?" that I think it fitting to digress for a moment to consider that question.

Poison gas has not been used in Europe for the following probable reasons:

1. To date, the campaigns have been rapid advances in a war of movement, whereas poison gas could be employed to the greatest advantage in a condition of stabilized warfare. Should a condition of stalemate result in trench warfare, as on the western front of World War I, we may see the employment of poison gas on a wide scale.
2. The fear of reprisal, particularly against civilians, may have been a deterrent. Airplanes on either side can spray mustard with terrible effect unless the civilian population is well tranied in protective measures. Knowledge that the people of Great Britain are all supplied with gas masks may have deterred Hitler from using toxic gases.
3. Conscientious scruples against the use of gas may have prevented its employment, but I believe that our enemies have given very little consideration to scruples or humanitarianism in the present conflict.

Poison gas has been used by the Japanese against the Chinese in the present conflict. Not on a large scale, to be sure, but sporadically, when they were hard pressed or apparently when they wished to experiment or try out some of their chemical munitions. Mustard, Lewisite, and chlorpicrin have all been reported as being used by the Japanese, while I have seen photographs of Chinese soldiers whose bodies bore large blisters caused by some vesicant agent.

The Italians used mustard with terrible effect against the barefooted and defenseless Ethiopians in 1936. It has frequently been reported that both sides in the present conflict have plentiful supplies of chemical agents on hand ready to retaliate should poison gas be used against them. No one can foresee when poison gas will be used, but it is not only possibile, but probable that a desperate Germany, with its highly developed chemical industry, may resort to the use of poison gas to stave off defeat. As to the Japanese, after Pearl Harbor, we can expect anything.

In this connection I want to correct a common misunderstanding. The United States is not a party to any treaty prohibiting the use of poison gas. The Geneva Protocol was not adopted by the United States Senate, so our armed forces are free to employ chemical agents should our Government decide to do so. Our armed forces are prepared to retaliate should poison gas be used against us.

RESEARCH AND TECHNICAL DEVELOPMENT

Constant research is essential for success in modern chemical industry but is of even greater importance in chemical warfare, because surprise is one of the fundamental principles of successful combat. Starting from scratch with feverish haste in 1915, research and development in chemical warfare have been carried on by all leading nations since that time. Nearly a hundred of the most prominent chemists at our leading colleges and universities are carrying on research problems in coöperation with the Chemical Warfare Service.

In general, the research problems may be grouped under three headings: (1) weapons, (2) munitions, (3) protective devices. To give you a better idea of the research problems, I can list the following as typical:

1. Improvements in the design of weapons.
2. Investigations of possible new agents.
3. Toxicity tests.
4. Improved methods for detecting chemical agents.
5. Studies in catalysts.
6. Production methods for agents.
7. Improvements in activated charcoal.
8. Absorbents and filtering materials.
9. Development of new decontaminants.
10. Physiological effects under field conditions.

The above list includes only a few of the hundreds of research problems being carried on by the Chemical Warfare Service.

Throughout the World War I, the lists of known organic and inorganic compounds were combed repeatedly by chemists of the highest caliber to find effective chemical warfare agents. It is interesting to note that all of the agents successful under field condi-

[1] Presented at the Fourth Summer Conference, New England Association of Chemistry Teachers, Durham, New Hampshire, August 13, 1942.

tions were compounds that had been known for many years. Research has been going on constantly since 1918. Possibly new compounds or new derivatives of well-known compounds have been developed, but no public announcements have been made. Should some new compound suddenly be employed in the battlefield, it is probable that our present protective equipment, possibly with slight modifications, will protect our troops against casualties. In chemical warfare, as in all types of combat, there is the continuous race between offense and defense, guns versus armor, fire power versus vulnerability, chemical agents versus protective devices. It is only by constant research and developmental work that our nation can be prepared for both offense and defense in chemical warfare —and this we are carrying on.

PRODUCTION

Our biggest task at present is production. Gas masks, protective clothing, chemical weapons, flame throwers, chemical agents, shells, grenades and smoke pots, as well as incendiary bombs, are needed in prodigious quantities for our rapidly expanding army. Through the foresight of the Service in preparing detailed procurement plans during peacetime, we were able to go into production rapidly as soon as the present emergency arose. We are very proud of the fact that we had a gas mask ready for each individual soldier as soon as he began his training. A part of our production program is being carried on at our arsenals, but the greater part is being produced by private industry working under contracts. All of the materials manufactured for the Army are subjected to close inspection throughout each stage in their processing to make sure that they conform to exact specifications. We have had excellent coöperation from the many industrial firms, without whose assistance our program would certainly have fallen down.

The manufacture of gas masks is not simple. Our service gas masks have about 100 individual parts, each of which must function perfectly or the value of the mask becomes nil. Rubber facepieces that are fully molded in one operation have been developed, while laminated, gas-resistant fabric can be used if rubber is not available. In addition to supplying our armed forces, the Chemical Warfare Service is producing gas masks for distribution to civilians by the Office of Civilian Defense. Production becomes a big task when you talk, not in thousands, but in millions. The problem is multiplied enormously when each item is made of many separate component parts.

Wars are not won by defensive measures, but rather by offensive operations. Hence we must produce large quantities of materials to be used, should occasion require. Here we enter a field of particular interest to the chemist. Besides smoke-producing chemicals, such as chlor-sulfonic acid and titanium tetrachloride, various toxic gases are required. Such items cannot be produced overnight; consequently, plants must be constructed and in operation long before the probable need for such munitions on the battlefield. While we purchase chlorine, acids, and various other commercial chemicals, the manufacture of chemical agents is for the most part carried on at chemical warfare arsenals.

There are several reasons why these agents are produced at our arsenals rather than by private industry. Chemical warfare agents form a highly specialized branch of chemistry which is non-industrial in nature. Industrial firms are not experienced in the manufacture of such agents nor are their plants equipped to carry on the various operations. Since the demand for chemical warfare agents exists only in time of war, it would not be profitable for a firm to devote time, money, and equipment or factory space to these products. In addition, the manufacture of chemical warfare agents inevitably involves certain industrial hazards which make it advisable to carry on the operations at Government arsenals where proper safeguards, dictated by years of experience, can be maintained.

While the manufacture of the common chemical warfare agents is described in such textbooks as Prentiss' "Chemicals in War," a brief outline of the method for making phosgene and mustard, two of the most important, may be of interest.

Phosgene, at ordinary temperatures and pressures, is a heavy colorless gas which condenses at 8°C. It was discovered by Davy in 1812; thus it was known for more than one hundred years prior to the first World War. Phosgene is prepared by the direct synthesis of chlorine and carbon monoxide with the aid of carbon as a catalyst. Carbon dioxide is first produced by burning coke. After cleaning and purifying, the carbon dioxide is stored in a holder. Oxygen is prepared from liquid air and is stored in a holder until required. Carbon monoxide is prepared by the reduction of carbon dioxide passing through a coke furnace. Sufficient oxygen is admitted to maintain the required temperature for the endothermic reaction.

Carbon monoxide and dry chlorine are admitted in proper proportions to a mixing chamber, from which the mixed gases are led to the catalyst chambers. After passing through the catalyst boxes the phosgene is condensed by means of refrigerating coils, while the uncondensed gases pass on to an absorption plant.

Phosgene is not difficult or expensive to make, and is an efficient lung irritant gas for war purposes. A considerable problem in the manufacture of this agent is the corrosion of steel and other metal parts by traces of hydrochloric acid vapor produced by hydrolysis.

Mustard, which is technically known as β,β-dichlordiethyl sulfide, is considered one of the most effective chemical warfare agents. Besides a powerful lung irritant action, mustard has a vesicant action on the skin, producing large water-filled blisters when the skin comes in contact with either the liquid or vapors. Mustard was first prepared in 1822 by Despretz, who prepared it by the action of ethylene on sulfur chloride.

It was later prepared by Guthrie in 1860, and also by Niemann in the same year. Meyer prepared mustard by the chlorination of thiodiglycol in 1886, and his method was used for the preparation of this agent by the Germans during the first World War. It is interesting to note that in spite of the enormous advances in chemistry, particularly in organic synthesis from the time of Kekulé, mustard, sometimes called the "king of war gases," was prepared more than a century ago.

The American method of producing mustard is to bubble ethylene gas through sulfur monochloride with careful control of temperature. The ethylene is produced from ethyl alcohol and is then scrubbed, dried, and stored until needed. Sulfur monochloride is produced by combining chlorine with sulfur. The reaction by which the mustard is produced, like so many others on paper, appears simple, but in large-scale production there are a number of production difficulties. During the last war, J. B. Conant, now President of Harvard College and head of the National Defense Research Council, worked out methods to overcome these production difficulties.

INSTRUCTION

The third major responsibility of the Chemical Warfare Service is to instruct all of the armed forces of the United States in chemical warfare. Our Service cannot instruct each separate unit, but does train officers and enlisted men of the Army, Navy, Marine Corps, and Coast Guard, who will in turn instruct the members of their units. This instruction is given at the Chemical Warfare School, which is located at Edgewood Arsenal in Maryland. The aim is to provide at least one graduate of the Chemical Warfare School for each of the larger components, such as a regiment, in order that our armed forces, down to the newest recruits, may be thoroughly grounded in defense against chemical attack. While chemical agents are capable of producing casualties, the greatest effect is psychological. Man's instinctive fear of something new and strange, such as "gas," can only be overcome by constant drill and thorough instruction. If we can teach the soldier to have confidence in his gas mask, and see that he knows how to wear it properly, we can overcome the psychological factor easily.

The present war has demonstrated that everybody—civilian as well as soldier—is subject to enemy attack. The devastating air raids which the Germans launched upon the English cities and towns have shown the need for training the civilian population in methods of defense against high explosives, incendiaries, and toxic gases. In order that our civilian population may be properly trained, the War Department Civilian Protection Schools have been organized under the control of the Chemical Warfare Service. Intensive ten-day courses are conducted for firemen, policemen, public utility employees, and other civic representatives at the five such schools now in operation, and two more will be organized shortly. These schools are located at Amherst College in Massachusetts, University of Maryland, University of Florida, Texas A. and M. College, and Stanford University in California, while the two new schools will be located in the Great Lakes Area and the Pacific Northwest. Just as in the instruction of our armed forces, the aim is to train leaders who will in turn train the citizens in their home communities. The importance of this work cannot be overemphasized. We may be subjected to enemy attack, particularly along the Atlantic and Pacific coastline, while the concentrated industrial area of Detroit would certainly be a prize target for our enemies. During 1937 and 1938, the British public displayed a considerable amount of apathy toward civil defense. The raids on London in December, 1940, caused widespread destruction. Later, raids in the spring of 1941 against the English Midlands cities produced only one-quarter as much damage per bomb dropped. This improved showing, from our point of view, resulted from the added impetus given to civil protection training. Let us hope that our nation will profit by the experience of Great Britain and give sufficient attention to civilian protection before the raids occur.

CHEMICAL TROOPS

The Chemical Warfare Service is also a combat arm of the Army, for we have chemical troops. These units include Weapons Companies, Decontamination Companies, Field Laboratory Companies, and Depot Companies. In addition, we have chemical companies assigned to duties in coöperation with the Air Forces. When the attack on Pearl Harbor occurred, one of our Chemical Companies was on the "alert" and used machine guns with telling effect upon the Japanese attackers.

The principal weapon of the Chemical Warfare Service is the 4.2-inch chemical mortar. This is a muzzle-loading, rifled weapon which nicely combines portability, rapidity of fire, and accuracy. Shells from this mortar can be fired a considerable distance. The filling of the shells can be white phosphorus, mustard, phosgene, or other toxic gases, as well as high explosive. The weapon was developed by Chemical Warfare Service officers and we are proud of its performance.

Livens projectors which fire a large watermelon-shaped projectile from a ground emplacement are valuable for laying down a heavy concentration of gas on a particular target. Chemical cylinders and smoke pots are other weapons of chemical troops.

The spraying of smoke or other agents from spray tanks fastened to the fuselage of an airplane is a method of dispersing our agents that has not been neglected.

Decontamination companies specialize in the decontamination of areas which have been gassed with persistent agents. Depot companies, as their name implies, are supply companies for issuing the various chemical warfare supplies to chemical troops and other branches of the field forces.

Of particular interest to chemists are our Field Laboratory Companies. These units operate in the field close to the front lines. The purpose of the Field Laboratory Company is to make analyses of chemical agents projected against our troops. The supplies and equipment can be quickly unpacked and set up. Bottled gas supplies fuel for Bunsen burners, while electric power is available from an accompanying truck generator to operate exhaust fans, electric furnaces, and other apparatus. It seems strange to see complete analyses performed in a tent that is set up in a wooded area to provide concealment from enemy planes, but I can testify that the personnel perform creditable work under these field conditions. The Field Laboratory Companies are the most highly educated companies in our Army. More than half of the personnel are graduate chemists, while Ph.D's are as common in these companies as corporals' chevrons in an infantry rifle company. By performing analyses on the spot, the Field Laboratory Companies are able to advise the Army Commander as to appropriate protective measures against new or surprise gases.

The Chemical Warfare Service has expanded at least as fast as other branches of our Army in the last two years. To take care of increased production, new Chemical Warfare Service arsenals have been constructed, while our original "mother arsenal" at Edgewood has been modernized and greatly expanded. Several plants have been constructed for manufacturing special chemicals for our Service, and a proving ground is under construction in Utah, where firing tests of our munitions can be carried on far from congested cities.

The incendiary bomb program of the Chemical Warfare Service is a tremendous undertaking and one of our main activities. Incendiaries have played a big part in the present conflict, and our program calls for the production of prodigious quantities, which should certainly send chills down the backs of our enemies. Newspaper accounts from Tokio indicate that Japan has already had a taste of what is to come.

Chemists, particularly, are interested in technical developments resulting from the war. After the conflict is over we shall be able to publish many interesting technical developments which cannot be discussed at present. The conservation of critical materials is a subject that is being carefully studied. New alloys to conserve tin are being developed, new bearing metals have been produced, while the Chemical Warfare Service is now using some plastic parts for gas masks instead of brass. The substitution of plastic parts has resulted in faster quantity production, fewer finishing operations, conservation of brass; and has also produced parts that are non-tarnishing, and require no paint or lacquer. The use of cellulose acetate lenses for gas masks instead of glass presents certain advantages. Laminated paper containers have been substituted for metal for packing cases, while Parkerizing has replaced cadmium plating for some metal components. In the manufacturing field, expensive glass-lined containers and reactors have been replaced by black iron vessels with corrosion-proof coatings of synthetic plastics. These are but a few of some of the technical developments which the war is bringing forth in our field.

WELCOME TO THE NEW ENGLAND ASSOCIATION OF CHEMISTRY TEACHERS

BEGINNING this month, the JOURNAL undertakes a new responsibility. An agreement has been made by which the *Report* of the New England Association of Chemistry Teachers is henceforth to be published in the pages of the JOURNAL.

The NEACT is by far the best-known and strongest chemistry teachers' association in the country. Its membership includes both college and secondary school teachers, and it has probably done more than any other one agency in establishing and promoting relations between these two groups. Its *Report* is now in its forty-fourth volume, has had a recognized place in the chemical literature, and has circulated widely, both in this country and abroad. We feel proud to be the means of continuing its existence in what we sincerely hope will be a wider sphere of usefulness.

We are sure that the arrangement will be of advantage both to the Association and to the JOURNAL. Members of the Association will not only receive their *Report* more frequently, but the rest of the JOURNAL as well. The JOURNAL, on the other hand, will have a wider circulation, especially in secondary school circles. We shall try to take advantage of this.

A new associate editor, Dr. Laurence S. Foster, of Brown University, has been nominated by the Association, and will be in charge of the *Report*.

Fountain pens and automatic pencils used up 2800 tons of steel last year, enough for 430,000 shells for 75-mm. field guns.

The Crystal Structure of Pond Ice Observed with Polaroid Spectacles

WILLIAM S. VON ARX *Brown University, Providence, Rhode Island*

NEARLY everyone has worn Polaroid spectacles at one time or another and noticed that the sky appears to change its brightness and color when the head is tilted from side to side. If a slab of ice cut from a pond is held up to the sky and examined while one's head is tilted to the angle at which the sky appears darkest, the ice crystals are revealed in colorful contrast.

This is a phenomenon of polarization. Certain objects appear strongly colored when examined in polarized light because of their anisotropism or property of "double refraction." Many natural crystals are anisotropic—calcite is exceptionally so—and most transparent substances will become so if they are subjected to mechanical strain. Within anisotropic substances, the velocity of light is different in different directions and for different frequencies of vibration and varying directions of approach. Therefore it commonly happens that several wave trains interfere destructively with each other while others interfere constructively. The destructive interference subtracts certain frequencies from the beam of light that finally gets through, with the result that it is colored. The depth of color (more precisely, the order of interference) is a direct function of the degree of anisotropism and thickness of the substance. In order to observe the interference color, the light entering the substance must be plane polarized, as from the sky, and the light emerging from the substance must be examined through a plane polarizing medium at right angles to the first. In this manner only the interference colors produced in a single plane are observed. Were it otherwise, the transmitted light would be practically white again due to the great variety of angles and velocities producing a confusion of colors simultaneously.

The sky looks dark when observed in the proper way through Polaroid spectacles because light from the sky is partly plane polarized. Since light which has passed through Polaroid lenses is also considerably plane polarized the amount of light reaching the eyes depends upon the angle between the two planes of polarization. When they are parallel, almost half of the light from the sky reaches the eyes. When they are perpendicular only the small unpolarized component of the sky light gets through. It can be shown by computations from Brewster's Law that sunlight, reflected from molecules of air overhead, is maximally polarized at an angle of very nearly 45°. Since the angle of reflection is equal to the angle of incidence, it is evident that one will observe maximal polarization in sky light when he looks at right angles to the incoming sunlight. A rule for observing the maximum dimming of the sky with Polaroid spectacles: point the top of the head at the sun while looking upward. The angle is frequently awkward, but in this position the plane of polarization of the sky light is at right angles to the plane of transmission of the Polaroid lenses; hence a minimum of light gets through the lenses. This is called the extinction position. Extinction in plane polarized monochromatic light is almost black. But light from the sky is composed of all wave lengths and not all of it is polarized in the same plane; hence the extinction is rather imperfect. Herepathite (tetraiodo quinine sulfate), the crystalline polarizing compound in Polaroid, is slightly less efficient in polarizing deep red and deep blue light than it is for the yellows and greens. Therefore the sky will appear rather a deep gray-blue at extinction, the gray being due to the unpolarized component of the sky light itself, and the blue cast a result of the less perfect action of Herepathite in polarizing blue light. It is this small degree of imperfection which makes the interference phenomena observed in this manner more than usually colorful and spectacular.

One other factor involved in this observational technic is the fact that light from the sky converges strongly upon the pupil of the eye. Therefore, if one examines a section of pond ice which is approximately plane-parallel, the light passing through the center of the sheet, when held squarely, travels a shorter path through the ice than the light entering obliquely. Recalling that the interference color is a function of the thickness of the section, it is evident that the colors in different parts of the section will vary. They will be arranged in concentric rings around the ray traveling the shortest path. Each successive ring is composed of the spectral colors, which grow fainter and more confused as the diameters of the rings increase. Across the rings of color, dark bands will be seen which correspond to the planes of polarization of the sky light and the Polaroid glasses. These dark bands are called isogyres, and by their means the crystallographer can tell the orientation of the crystal he is observing.

In pond ice, for instance, there are two distinct modes of orientation. When the pond first freezes over, the ice crystals grow from the colder water near the shore into the cooling water in the middle. Ice crystals grow with their principal axes parallel with the temperature gradient. Therefore the first skin of ice will be composed of horizontal crystals in roughly convergent array toward the center of the pond. A section of this early

ice held up to the sky for examination in polarized light will show colored rings but the isogyres will move across each of the crystals in the section very swiftly as it is rotated on the line of sight. This is the flash figure which indicates that the principal crystallographic axis is perpendicular to the line of sight. If one waits for the later growth of ice which develops straight downward into the cooling interior of the pond after it is completely covered with crystals, and then cuts a section, he will observe in each of the crystals a pair of isogyres resembling the Maltese cross. These isogyres remain relatively stationary when the section is rotated.

The pattern is the optic axis figure which indicates that one is looking along the principal axis of the hexagonal ice crystal. Upon closer scrutiny of the crystal sections one may see the undulatory extinction pattern which is caused either by strains due to unequalized pressure within the ice or to discontinuous crystallization at the time the ice was formed. Large "single crystals" of ice may in this way be found in pond ice when the conditions of crystallization have been favorable.

A Semimicro Scheme of Qualitative Analysis[1]

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WITHIN the last several years many attempts have been made to devise a scheme of inorganic qualitative analysis for the common metal ions without the use of hydrogen sulfide. To be sure, hydrogen sulfide has its objectionable features, such as repulsive odor, toxicity, and difficulty of administration. Most of these schemes involve separations depending on the insolubility of the oxides, chlorides, sulfates, hydroxides and carbonates of the metal ions. The difference in the solubility products of these various compounds is not sufficient to give complete and clear-cut separations. Consequently it is unlikely that as far as logical and practical chemistry is concerned the classical scheme, with the use of hydrogen sulfide, will be displaced.

Meanwhile, many new reagents for the identification of various metal ions have been developed or discovered. Strange as it may seem, no extensive work has been done to adapt these new reagents to a complete systematic macro or semimicro qualitative scheme of inorganic analysis. They have been used rather extensively for micro analysis, chiefly because of their sensitivity, and for the so-called "spot tests." Some semimicro schemes have been proposed, but these involve, for the most part, the same type of reactions and identification tests as those employed in the classical macro scheme, the only difference being that the operations are carried out on smaller quantities.

Realizing that from a scientific and pedagogical standpoint the development of a method superior to the sulfide separation is quite improbable, there appear to be two ways by which this method may be improved very considerably. The first is to make use of the many specific and sensitive reagents now at hand, and thereby enable us to detect smaller quantities of the various metal ions. This would eliminate many of the slow filtrations and tedious separations. The second is to overcome the objections to the use of hydrogen sulfide. A semimicro scheme of qualitative analysis to achieve this goal is here described.

In the present investigation a number of organic compounds are used to detect certain of the metal ions. So many compounds have been proposed as qualitative reagents that the choice of the most suitable ones has been difficult. Those selected were chosen on the following basis:

1. That the method of making the test be a simple one.
2. That the reagent be as specific as possible and still meet other requirements, thereby making it possible to test for a metal ion in a solution containing several others.
3. That the reagent produce a very striking change in the appearance of the solution, for example a drastic change in color, or the formation of a heavy precipitate.
4. That the conditions of the test not be too limited.
5. That the cost of the reagent not be prohibitive.
6. That the reagent be stable in solution.

THE SCHEME

Aside from the fact that magnesium is placed in its natural group with barium, strontium, and calcium in Group IV, the groups are the same as those in the common scheme of analysis. The precipitating conditions and reagents are the same as in the usual scheme except that in Group IV the metal ions are precipitated as phosphates from an alkaline solution.

Working on a semimicro scale means that the apparatus must be selected to suit the scale. For all separations of solids and liquids the centrifuge is employed; centrifuge tubes of 3-ml. capacity and 75-mm. test tubes are used throughout. For transferring liquids, droppers of 1.25-ml. capacity are used, and for very small amounts of liquids a semimicro capillary pipet with a capacity of 0.2-ml. is quite serviceable. For all

[1] Taken from a thesis presented by J. A. Southern to the faculty of the University of North Carolina in partial fulfilment of the requirements for the Doctor's Degree.

[2] At present with the Chemical Warfare Service, New York, N. Y.

heating and warming processes the tubes are suspended in a beaker of boiling water, being held in place by a cover of sheet copper with holes large enough to support the tubes. One-ml. pipets graduated in 0.1 ml., are used for accurate measurement of liquids.

The systematic analysis is carried out as follows:

GROUP I

Procedure 1. Precipitation of Group I. To 5 drops of the unknown, contained in a centrifuge tube, add 1 drop of 3 *N* HCl. Centrifuge until clear. Test for complete precipitation by adding 1 drop HCl. The *precipitate* is composed of the chlorides of the Group I metals, and after removal of the centrifugate is washed with 2 drops of water and analyzed according to Procedure 2. The *centrifugate* containing the metals of Groups II to V is preserved for Group II (Procedure 3).

Procedure 2. Treat the white precipitate Procedure 1 with 2 or 3 drops of water, heat in the water bath one minute and centrifuge.

	Residue may contain AgCl and Hg_2Cl_2. Treat with 2 drops of 3 *N* NH_4OH and centrifuge. Remove the centrifugate.
Centrifugate may contain $PbCl_2$. Divide into two portions. 1. To the first, add an equal portion of 2.5 *N* NaAc and heat in water bath. When hot, add 10 drops of bromine water and continue to heat for 1 to 2 minutes. Centrifuge and pour off the liquid. Wash the residue with 10 drops of water. Wash again with 5 drops of water to which 5 drops of 3 *N* NH_4OH have been added. Discard the wash water. Then add 5 drops of benzidine solution. The formation of a purplish blue solution shows the presence of LEAD. 2. To the second portion add K_2CrO_4. A yellow precipitate shows the presence of LEAD	*Centrifugate* may contain AgCl. Make acid with HNO_3. The formation of white precipitate shows the presence of SILVER.	*Residue* may contain Hg. Wash with 10 drops of water. Discard the wash water. Dissolve in 1 drop of conc. HNO_3. Dilute to 1 ml. with water and divide into two portions. 1. Make first portion neutral with 2 *N* KOH, add 10 drops of 2.5 *N* NaAc and 2 drops of diphenylcarbazid solution. The formation of a deep purple color shows the presence of MERCURY. 2. To second portion add $SnCl_2$. A white precipitate shows the presence of MERCURY.

GROUP II

Procedure 3. Precipitation of Group II. To the centrifugate from Group I (Procedure 1) add one drop of 3 *N* HNO_3. Heat in the water bath three minutes. Add 3 *N* NH_4OH until the solution is barely alkaline, then add 3 *N* HCl dropwise until the solution is barely acid, concentrate by boiling to 1 ml. if necessary, then add 1 drop of concentrated HCl. Heat to boiling, place in water bath and pass in H_2S slowly for three minutes. Cool, dilute to 2 ml. and again pass in H_2S for thirty seconds. A white precipitate is sulfur, not Group II. Centrifuge and test for complete precipitation. When precipitation is complete, centrifuge and remove the centrifugate.

Centrifugate contains Groups III to V. Heat in water bath for several minutes, or until all H_2S has been expelled and preserve for Group III (Procedure 7).

Residue consists of sulfides of metals of Group II and is treated according to Procedure 4.

Procedure 4. Separation of Groups IIA and IIB. The sulfide precipitate from Procedure 3 is washed with 10 drops of water containing a drop of saturated NH_4Cl. The wash water is discarded. Treat the residue with 4 drops of 2 *N* KOH and heat in water bath for about two minutes and centrifuge. Remove centrifugate. Wash the residue with 4 drops of KOH and centrifuge. Remove wash solution and add to centrifugate.

Centrifugate contains the thio and oxythio salts of Group IIB (as Sb and Sn) and is analyzed according to Procedure 6.

Residue contains the sulfides of Group IIA (Hg, Pb, Bi, Cu, and Cd) and is analyzed according to Procedure 5.

Procedure 5. Analysis of Group IIA. Wash the residue from Procedure 4 with 10 drops of water, discard the washings and then add 5 drops of 3 *N* HNO_3; warm for two minutes and centrifuge.

	Centrifugate contains nitrates of Cu, Cd, Bi, and Pb. Add 5 drops of concentrated NH_4OH and centrifuge. Wash with 5 drops of water and add washings to the centrifugate.		
Residue may be black HgS or white $2HgS{\cdot}Hg(NO_3)_2$. Dissolve in 1 or 2 drops of aqua regia by heating. Divide into two portions. 1. To first portion add 2 drops of diphenylcarbazide solution. Now let 2 *N* KOH run down the side of the inclined tube. A bluish purple color shows the presence of MERCURY. 2. To second portion add $SnCl_2$. A white precipitate shows the presence of MERCURY.	*Centrifugate* contains Cu and Cd as complex ions. Divide into two portions. 1. To first portion add 2 drops of benzoin oxime solution. The formation of a greenish precipitate shows the presence of COPPER. 2. To second portion add a drop of KCN. Divide into two portions. (*a*) To first portion add 6 drops of KOH and 2 drops of "Cadion" and then 6 to 10 drops of formaldehyde. The separation of a reddish colored precipitate shows the presence of CADMIUM. (*b*) To second portion add H_2S. A greenish yellow precipitate shows the presence of CADMIUM.	*Residue* consists of $Pb(OH)_2$ and $Bi(OH)_3$. Add 2 drops of KOH, warm and centrifuge. Remove the centrifugate.	
		Residue may be $Bi(OH)_3$. Wash with water. Discard washings. Dissolve in 1 drop of 3 *N* HCl. Add 10 drops of cinchonine-KI reagent. An intensely orange colored precipitate shows the presence of BISMUTH.	*Centrifugate* may contain Pb as $PbO_2^=$. Make barely acid with 3 *N* HAc. Divide into two portions. 1. To first portion add 10 drops of 2.5 *N* NaAc. Test for LEAD as in Group I. 2. To second portion add K_2CrO_4. A yellow precipitate shows the presence of LEAD.

Procedure 6. Analysis of Group IIB. To the centrifugate (Procedure 4) containing the metals of Group IIB, add dilute HCl until neutral, centrifuge and discard the centrifugate, then add 6 drops of concentrated HCl, stir, and centrifuge.

Centrifugate contains Sb and Sn as chlorides. Divide into two portions. 1. To the first portion add 0.05 g. $NaNO_2$ slowly, then 2 drops of 3 *N* HCl. Finally, to 2 drops of solution on a spot plate add 2 drops of Rhodamine B solution. A purplish color shows the presence of ANTIMONY. 2. To second portion add an equal volume of 3 *N* HCl. Add a piece of magnesium ribbon about 2 cm. long and allow to stand until the magnesium is dissolved. Add 5 drops of water and 1 drop of phosphomolybdic acid. The formation of a blue color shows the presence of TIN.	*Residue* may contain As. Dissolve in a drop of concentrated HNO_3 by heating and add 5 drops of $(NH_4)_2MoO_4$. Heat 1 to 2 minutes, dilute to 3 ml. and cool. Add 2 drops of $SnCl_2$. The formation of a deep blue color shows the presence of ARSENIC.

GROUP III

Procedure 7. Precipitation of Group III. To the centrifugate from Group II (Procedure 3) add 2 drops of NH_4Cl and concentrated NH_4OH until distinctly alkaline (0.1 ml. should be sufficient). Pass in H_2S until precipitation is complete and centrifuge.

Precipitate is to be analyzed according to Procedure 8.

Centrifugate contains Groups IV and V. Make acid with acetic acid and heat in the water bath until analysis of Group III is completed. Centrifuge and discard any precipitate which may form while the solution is being heated. Save the clear, colorless centrifugate for analysis of Groups IV and V (Procedure 11).

Procedure 8. Analysis of Group III. Wash the precipitate from Procedure 7, consisting of the hydroxides of Cr and Al and sulfides of Zn, Fe, Co, Ni, and Mn, with 10 drops of water to which 2 drops of 4 *N* NH_4Cl have been added, and dissolve by adding first 1 or 2 drops of concentrated HCl and heating in water bath for one minute, then add a small drop of concentrated HNO_3. Continue heating until the solution is clear. Make the solution alkaline with KOH and slowly add 0.1 to 0.2 gm. Na_2O_2, heat for one minute in water bath and centrifuge.

Centrifugate contains Cr, Zn, and Al and is analyzed according to Procedure 9.

Precipitate contains Fe, Co, Ni, and Mn as hydroxides or hydrated oxides and is analyzed according to Procedure 10.

Procedure 9. Centrifugate from Procedure 8 contains AlO_2^-, $CrO_4^=$ and $ZnO_2^=$. Divide into three portions.

1st Portion	2nd Portion	3rd Portion
Test for Cr by heating for one minute and adding an equal volume of water, and making barely acid with H_2SO_4. Then add 2 drops of diphenylcarbazide solution. The formation of a reddish violet color shows the presence of CHROMIUM.	Test for Al as follows: Acidify with 3 *N* HAc. Add 2 drops of "Aluminon." Make barely alkaline with NH_4OH and centrifuge. The separation of a red precipitate shows the presence of ALUMINUM.	Test for Zn as follows: Add an equal volume of water and 10 drops of dithizone solution. *Shake vigorously.* A change in the color of the aqueous layer to pink, not brown, shows the presence of ZINC.

Procedure 10. The precipitate from Procedure 8 consists of the hydroxides and hydrated oxides of Fe, Co, Ni, and Mn. Wash with 10 drops of water. Dissolve in concentrated HCl. Dilute and divide into four portions.

1st Portion	2nd Portion	3rd Portion	4th Portion
Test for Mn as follows: Add 2 drops of saturated $NaKC_4H_4O_6$. Make alkaline with 2 *N* KOH and shake for one minute. Add 1 drop of benzidine. A blue color shows the presence of MANGANESE	Test for Fe by adding 1 drop of 3 *N* NH_4CNS. The formation of a deep wine red color shows the presence of IRON.	Test for Co as follows: Add solid Na_2F_2 until the solution is colorless. Now add 2 drops of NH_4CNS. The formation of a blue color shows the presence of COBALT.	Test for Ni as follows: Add solid Na_2F_2 until the solution is colorless. Make alkaline with 3 *N* NH_4OH, then add 2 drops of dimethylglyoxime. The formation of a red precipitate shows the presence of NICKEL.

GROUP IV

Procedure 11. Detection of Mg. The clear, colorless centrifugate from Group III (Procedure 7) should have a volume of not greater than 1.5 ml. If larger, heat in the water bath until the volume is reduced. Discard any residue which may settle out. Take one-fifth of the solution (save the rest for Procedure 12) and test for Mg as follows: add 2 drops of the Mg reagent (*p*-nitrobenzeneazo-*β*-naphthol) and then 4 drops of 2 *N* KOH. Shake well and centrifuge. The formation of a purple precipitate shows the presence of MAGNESIUM.

Procedure 12. Precipitation of Group IV. To the remainder of the solution, saved from Procedure 11, add NH_4OH until alkaline, then 0.6 *N* $(NH_4)_2HPO_4$ until precipitation is complete and centrifuge.

Precipitate contains phosphates of Ba, Sr, Ca, and Mg and is analyzed according to Procedure 13.

Centrifugate contains Group V and is analyzed according to Procedure 14.

Procedure 13. Wash the precipitate from Procedure 12 with 10 drops of water containing a small drop of 3 *N* NH_4OH. Dissolve in a drop of 3 *N* HAc. Add one drop of K_2CrO_4, allow to stand for one minute and centrifuge. Wash the precipitate with a drop of water, adding the washings to the centrifugate.

Precipitate may be $BaCrO_4$ and indicates the presence of BARIUM. Confirm with flame test.	*Centrifugate* contains Sr, Ca, and Mg as acetates. Make alkaline with 3 *N* NH_4OH and centrifuge. Discard the centrifugate. *Precipitate* consists of the phosphates of Sr, Ca, and Mg. Wash with 3 drops of water, discarding the washings. Dissolve the precipitate in a small drop of 3 *N* HAc and add an equal volume of water. Add a small drop of 2 *N* $(NH_4)_2SO_4$ and centrifuge.		
	Precipitate may be $SrSO_4$ and indicates the presence of STRONTIUM. Confirm with flame test.	*Centrifugate* contains $CaSO_4$ and $MgSO_4$. Add a small drop of $(NH_4)_2C_2O_4$ and centrituge.	
		Precipitate may be CaC_2O_4 and indicates the presence of CALCIUM. Confirm with flame test.	*Centrifugate* contains Mg but is to be discarded, as Mg was detected in Procedure 11.

GROUP V

Procedure 14. Removal of $PO_4^=$ ions. To the centrifugate from Group IV (Procedure 12) add $Mg(NO_3)_2$ until precipitation of $MgNH_4PO_4$ is complete and centrifuge.

Precipitate is discarded.

Centrifugate contains Na, K, and NH_4 salts and is analyzed according to Procedure 15.

Procedure 15. Removal of ammonium salts. Place the centrifugate from Procedure 14 in a small crucible, add 4 drops of 3 *N* H_2SO_4 and evaporate to dryness. Continue to heat just below red heat until all fumes cease to be evolved. Treat the *residue* as directed in Procedure 17.

Procedure 16. Test for the ammonium ion. Place a drop of the original unknown in a small test tube and add 2 drops of 2 *N* KOH. At the mouth of the tube hold a piece of filter paper moistened with Nessler's reagent, or a piece of moistened red litmus paper. Warm the solution. The Nessler's reagent will

turn yellow to brown, depending on the amount of NH_4^+ present. The litmus will turn blue, which shows the presence of NH_4^+.

Procedure 17. Test for Na and K. Cool the residue from Procedure 15 and add a drop of water and make flame test for sodium and potassium. To further confirm the presence of these metals add another drop of water and proceed as follows:

Test for Sodium	Test for Potassium
Take 1 drop of solution from crucible and put in centrifuge tube. Add 2 drops of uranyl acetate reagent for sodium, allow to stand for one to two minutes and centrifuge. A greenish yellow precipitate shows the presence of SODIUM.	To solution remaining in crucible add 1 drop of dipicrylamine solution. Allow to stand for several minutes. The formation of orange-red crystals and fading of the color of the solution from orange-red to yellow shows the presence of POTASSIUM.

CONCLUSION

The proposed scheme has been used with approximately 1200 students taking qualitative analysis. It is being used at the University of North Carolina and four other colleges in the mimeograph form and the results have been most satisfactory. Students have had fewer "repeats" with unknowns using the scheme than with schemes previously used. The proposed scheme has a good pedagogical value and observations have led us to feel that it is an improvement over the macro and semimicro schemes to be found in the literature and textbooks in the following respects: (*a*) ease of manipulation (*b*) accuracy of results (*c*) economy of materials and (*d*) a saving of time due to the reduction of the number of steps required to carry out the procedure. We have completed an analysis of a solution containing all metal ions in less than two hours.

The Industry of Solid Carbon Dioxide

D. H. KILLEFFER *New York City*

SUPPOSE you had gone into the corner grocery store today to buy the things for dinner. You wanted roast beef, and fresh green peas, even though the season for peas was past. You wanted strawberries for dessert, too, and cream to go on them. Walking home, having gotten everything you ordered, you stopped at an ice cream truck, and nibbled reflectively on ice cream the rest of the way home.

You probably weren't reflecting on your good fortune in being able to get fresh food, in season or out, and whether or not the food can be grown in your neighborhood.

But it is good fortune, and it is due, in a large part, to solid carbon dioxide. For solid carbon dioxide has become an important modern refrigerant. It fills the bunkers in railroad cars transporting meat from Chicago, or fish from the coasts. It preserves the low temperatures of frozen foods. It protects perishable fruits in transit. By being compact, it permits the ice cream salesman to increase his market, to use a larger amount of the space in his truck for the ice cream than for the ice necessary to keep it cold. No more heavy, water-logged containers, dripping corrosive brine wherever they stand.

But solid carbon dioxide was not always of such importance in daily living; solid carbon dioxide had to wait 91 years to find work. From its preparation for the first time in a French laboratory in 1835, until American ingenuity gave it a commercially useful job to do in 1926, this intensely cold material appeared only in laboratory demonstrations and occasionally was used by physicians to remove warts by freezing. In 1926 it appeared on the market as a remarkable special refrigerant, and since then has become an important commodity. In 1939, forty-four plants in the United States produced some 178 thousand tons of dry ice. The greater part of this production is used for refrigeration, principally of perishables in transit, and the remainder is a convenient form for transporting carbon dioxide for carbonating beverages.

Because of its picturesque properties and its phenomenal growth, the solid carbon dioxide industry has attracted numerous inventors and investigators who have sought to improve dry ice manufacture, and widen its usefulness. These inventors and investigators have taken out hundreds of patents on their processes; this despite the fact that solid carbon dioxide itself, first made by Thilorier in 1835,[1] is not subject to patent.

The properties of solid carbon dioxide are sufficiently unusual to suggest a host of applications. Most important of these is the fact that it sublimes directly from the solid to the gaseous state at pressures below 5.11 atmospheres, instead of passing from solid to liquid and then to gaseous state, as most substances do. At one atmosphere (1 atmosphere = 760 mm. mercury pressure), the temperature of sublimation is −78.5°C. (−109.3°F.), 141°F. colder than water ice. This temperature is far below any required in industry, except in the most unusual cases. Liquid air is, of course, much colder, but also much more expensive.

The gas evolved during sublimation has many valuable properties, since it is extremely pure, anhy-

[1] THILORIER, *Ann. de Chimie et Phys.*, **60**, 432 (1835).

drous carbon dioxide. Although it is commonly considered to be an inert gas, it is only so with respect to oxidation and reduction at ordinary temperatures. At very high temperatures it becomes an oxidizing agent, and will react with the active metals. In the presence of moisture it produces a distinct acid reaction. Living things, both animal and vegetable, are profoundly affected by carbon dioxide. Since the molecular weight of carbon dioxide is greater than that of oxygen or nitrogen, air and moisture diffuse into it, sometimes with actually measurable pressure. Rubber is highly permeable to carbon dioxide; so much so that a toy balloon, partially inflated with air, will actually blow itself up to the bursting point when it is immersed in an atmosphere of pure carbon dioxide. Another especially important characteristic of the gas is its relatively high insulating power against heat. For practical purposes, a quiescent layer of dry carbon dioxide (confined to prevent convection, which is the transmission of heat by means of currents) has about 50 per cent more insulating value than a similar layer of dry air. Part of this greater insulating capacity is imparted to fibrous insulants in which carbon dioxide replaces air. These characteristics of the evolved gas are often overlooked, but do play vital roles in the utilization of solid carbon dioxide.

Solid carbon dioxide has no true melting point in the ordinary sense, except under substantial pressure, but sublimes directly to the gas. This fact has been of the utmost significance in its commercial use for refrigeration. Its effect has been twofold; the latent heat of sublimation, equivalent to the latent heats of fusion and of vaporization, is relatively high; thus permitting the storage of a large amount of "cooling" in small weight and bulk. And, since no liquid phase is involved in sublimation, it has been practicable to build useful refrigerated containers, or trucks, or freight cars, from light-weight materials which would be impossible to use if carbon dioxide melted to a liquid. No drains are necessary, and there is no liquid to corrode the containing vessel.

PROCESS OF MANUFACTURE

Manufacture of solid carbon dioxide as normally practiced consists of three essential steps:

1. The concentration and purification of the gas.
2. The liquefaction of the gas.
3. Solidification.

The details of these steps are, of course, modified by circumstances. The concentration and purification of the gas, particularly, depend on the type of raw material available. Some sources of carbon dioxide yield substantially pure gas (95 per cent or more), containing no nitrogen, oxygen, or other "permanent" gases. In other sources, impurities of this character predominate. In the first category are the gases produced in alcoholic fermentation and as by-products in several chemical operations, and the gas from certain natural wells. In the second group are the gases from combustion, from lime and cement kilns, and from other natural wells and springs of high "permanent" gas content. Gases having relatively high concentrations of carbon dioxide are treated to *remove the impurities* from the carbon dioxide. When there are more "permanent" gases than carbon dioxide present in the mixture, however, the carbon dioxide is removed *from* the impurities.

Fermentation gas, in which the percentage of carbon dioxide is high, may be washed successively with water and with sulfuric acid and then passed, under pressure, over activated carbon or silica gel to remove residual odors. Other treatments with chemical reagents (oxidizing agents and adsorbents) have been suggested to remove odorous impurities. The deodorized gas is then compressed to the liquid form.

A natural gas which is chiefly carbon dioxide may be similarly purified by treatment with adsorbents and solvents to remove impurities, which are usually hydrocarbons, and then finally liquefied.

These processes are relatively simple and inexpensive when compared with the concentration of carbon dioxide from dilute mixtures, such as flue gas. In such cases advantage is taken of the greater chemical reactivity and solubility of carbon dioxide as compared with nitrogen by forming loosely bound chemical compounds with an alkaline material, such as sodium carbonate. The resulting bicarbonate is subsequently decomposed by heating to yield pure carbon dioxide.

Once the gas is separated in pure form, its liquefaction is accomplished by compression and cooling, the pressure ultimately being raised to about 70 atmospheres. This liquid carbon dioxide may be the commercial product marketed in cylinders, or it may be the raw material for solidification. If it is to be solidified, the

Courtesy of Pure Carbonic, Incorporated

BLOCKS OF COLD

Every seven minutes a 220-pound block of solid carbon dioxide drops out of this huge machine

Courtesy of Industrial and Engineering Chemistry
SAWING BLOCKS OF SOLID CARBON DIOXIDE

next step is to cool it still further, and finally, by its own evaporation, to solidify part of it.

The conversion of the liquid at 20°C. to the solid at −78.5°C. may be accomplished by three different methods. The entire cooling may be done by an external refrigerating system, in which case an ammonia refrigerating machine may be used to cool the warm liquid carbon dioxide to such a low temperature that the liquid freezes. Since this is usually not economical or practical, a second method is often used, in which an ammonia system cools the liquid only part way to the freezing point, and then part of the cool liquid is evaporated very rapidly, which cools the rest of the liquid to the freezing point to finish the process. A third method involves evaporating part of the liquid carbon dioxide in several stages and ultimately solidifying the rest of it by the cooling effect of its own evaporation.

The final result, whichever method is employed, is the production, by compression, of a solid, glassy cube, 10 inches on each edge and weighing approximately 50 pounds. Frequently the press in which the blocks are formed make pieces 10 × 20 × 20 inches which are sawed into four cubes on a band saw. The blocks are individually wrapped in paper bags to keep air away from them, and are stored or shipped in heavily insulated containers.

The major use of the solid carbon dioxide produced in the United States is in the refrigeration industries, which consume 85 per cent of the total. Ten per cent more of the total production is used for various other purposes, of which the most important are: the production of atmospheres of pure dry carbon dioxide for chemical or other uses; as a substitute for liquid air in condensation traps of high-vacuum systems; in preventing the self-hardening of aluminum alloy parts after heat treatment; in cooling cores for shrink fits; and in the finishing of molded rubber articles from which the mold fins can be readily broken when cooled to dry ice temperatures. The remaining five per cent provides a convenient method for supplying liquid carbon dioxide when it is needed.

The 85 per cent used in the refrigerating industries splits up as follows: 81 per cent is used in the refrigeration of ice cream, principally in trucks and transportation packages; the refrigeration of meat and fish in trucks, and to a smaller extent in railroad cars, accounts for another 14 per cent; the industry of frozen foods makes up the remaining five per cent. Dry ice, by reason of its low temperature, is the ideal refrigerant for the storage and shipment of ice cream. The meat packer is particularly interested in solid carbon dioxide refrigeration because the off-coming gas effects a definite improvement in the final condition of his product. Eggs in cold storage are maintained in a fresh condition for a much longer period if the atmosphere of the storage space is supplied with regulated amounts of carbon dioxide. The gas also benefits cut flowers.

Despite its advantages, the use of dry ice as a refrigerant seems likely to be limited by its cost. In mobile refrigeration of trucks and freight cars, the weight reduction which it makes possible yields a return great enough to offset its high cost. And, where extremely low temperatures are needed, it has a cost advantage over most refrigeration systems. Nevertheless, the cost of solid carbon dioxide is about ten times that of water ice in large wholesale quantities, and it produces only about twice as much refrigerative effect. Dollar for dollar, ice is five times as effective as dry ice. Although this difference may be considerably reduced in the future, it seems unlikely that the cost of cooling effect in the two materials will ever reach a par.

Improved methods of application may possibly widen the field of use of dry ice by increasing its efficiency. Costs may also be lowered by improved methods of manufacture. On the other hand, low temperature refrigeration of loads in transit will continue for some time to be the mainstay of the industry.

NEW DATES AND LOCATION FOR THE NATIONAL CHEMICAL EXHIBITION

ANNOUNCEMENT is made by the Chicago Section of the American Chemical Society of a change in dates and location for its second National Chemical Exposition, owing to the acquisition of the Stevens Hotel in Chicago by the United States Army.

The exposition and conference have been transferred to the Sherman Hotel, conveniently located in Chicago's business center at Clark and Randolph Streets, and will take place from November 24 to 29, instead of a week earlier as originally planned.

The Chemistry of Group I: Silver, Mercury, Lead

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Department of Chemistry, The Ohio State University

A NEED HAS been felt for a summary of the results of recent research that gives new information about the chemical facts and principles involved in qualitative analysis. It is the purpose of this paper to present such a summary of Group I.

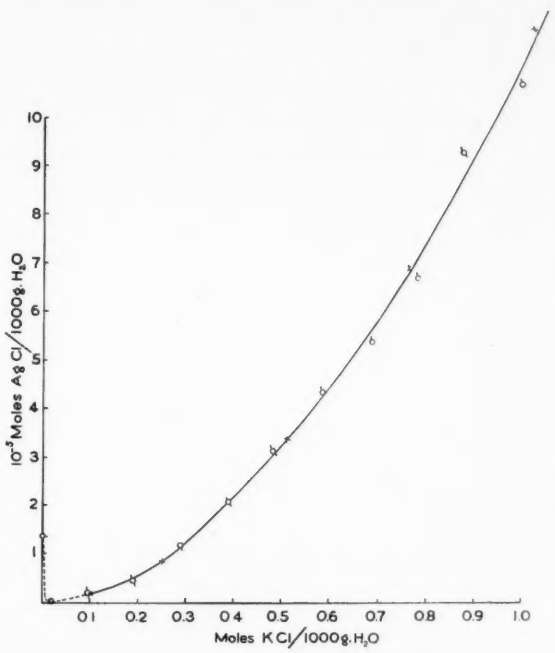

FIGURE 1

A classical procedure for the analysis of Group I is that in which the ions are precipitated as chlorides from which lead chloride is removed with hot water and silver chloride with ammonium hydroxide; reactions related to that procedure are discussed here.

EVIDENCE FOR MERCUROUS CHLORIDE AS THE POLYMER, Hg_2Cl_2

This formula, Hg_2Cl_2, for mercurous chloride is indicated in the solid by X-ray data (*1*) (which indicate the presence of the linear polymer Cl—Hg·Hg—Cl); in solution by conductivity and transport experiments (*2*); and in the gaseous state by vapor density. The molecular species present in the gas are uncertain above 250° but it is concluded from magnetic measurements (*3*) that none of the HgCl species is present in the vapor state. Mercurous chloride is diamagnetic in the solid, liquid, and gaseous state; if the molecular species were the monomer (HgCl) it would have an odd number of electrons and be paramagnetic.

The arrangement of the atoms in the crystal also indicates that we may expect Hg_2Cl_2 to be only very slightly dissociated (ionized) in solution.

THE EFFECT OF THE ADDITION OF CHLORIDE IONS ON THE CONCENTRATION OF SILVER, MERCURY, AND LEAD IONS

This effect is shown by means of the data on the solubility of silver chloride, mercurous chloride, and lead chloride in the presence of various amounts of chloride ions (*4*, *5*, *6*). (See Tables 1, 2, and 3 and Figures 1, 2, and 3.) With silver chloride and lead chloride the common ion effect is very pronounced. The minimum concentrations of lead and silver in solution appear to be at 0.8 and 0.01 molal added chloride, respectively, but the minimum concentration of mercurous mercury is at 0 molal concentration of added chloride which results from the fact that mercurous chloride, Hg_2Cl_2, has only a very low degree of dissociation (ionization), hence the common ion has little or no effect. (*Note:* The *molal* concentration is the number of moles of solute in 1000 grams of water.)

Further study of Figures 1, 2, and 3 indicates that the concentrations of silver, mercury, and lead all increase as the added chloride ion concentration increases above 0.01 molar, 0.0 molar, and 0.8 molar, respectively. This is interpreted to indicate that

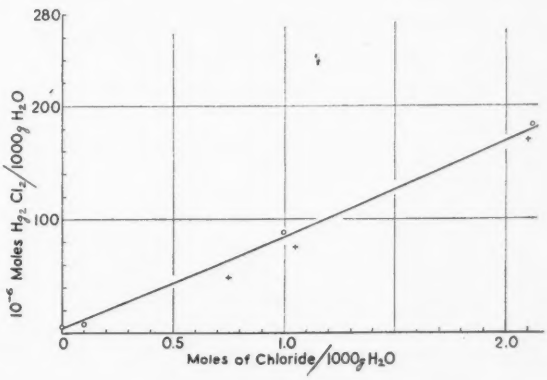

FIGURE 2

complex halide ions are formed of the character $PbCl_3^-$, $AgCl_3^=$, and $Hg_2Cl_3^-$. One method of determining the character of such ion species is to calculate the equilibrium constants for the probable reactions involved. That reaction which yields a true constant is chosen as the most probable one. Another method is to rearrange the expression for the equilibrium constant in such a manner that the number of ions reacting is indicated by a graphical method. This is illustrated in the determination of the ion species present in silver

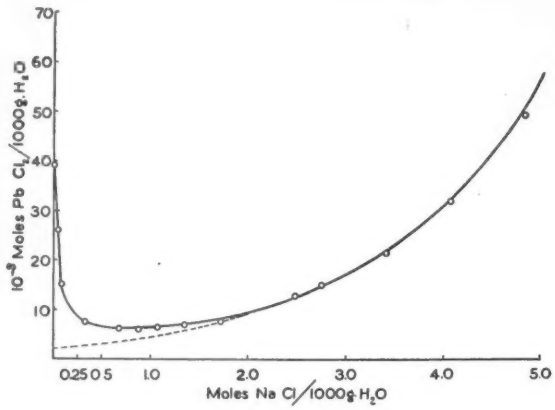

FIGURE 3

chloride and mercurous chloride solutions containing large amounts of added chloride ion.

(*a*) *The Complex Ions of Silver Chloride.* The general equation that represents the increased solubility of silver chloride is

$$AgCl_{(s)} + nCl^- = AgCl_{n+1}^{-n} \text{ (Complex ion)} \qquad (1)$$

$$K_1 = \frac{a_{AgCl_{n+1}^{-n}}}{a_{Cl}^{-n}}$$

where a = activity = molality × activity coefficient = $m\gamma$ and

$$\log K_1 = \log a_{AgCl_{n+1}^{-n}} - n \log a_{Cl^-} \qquad (3)$$

$$\log a_{AgCl_{n+1}^{-n}} = n \log a_{Cl^-} + \log K_1 \qquad (4)$$

If one plots $\log a_{AgCl_{n+1}^{-n}}$ against $\log a_{Cl^-}$ the slope will be n and the intercept $\log K_1$; this is shown in Figure 4. The value of $n = 2$ obtained from the slope indicates that the chief reaction in this range must certainly be

$$AgCl_{(s)} + 2Cl^- = AgCl_3^= \text{ and } K_s = 1.3 \times 10^{-4} \qquad (5)$$

Note: In this determination $\log m_{Cl^-}$ is actually plotted against $\log m_{Ag}$ complex. This is justified if $n = 1$ for in that case $\gamma_{AgCl_2}/\gamma_{Cl^-} = 1$; it can also be justified if $n = 2$ for, here, according to the Debye-Hückel relationship, $-\log \gamma = CZ_1Z_2\sqrt{\mu}$, it can be shown that at a given concentration (value of μ): $\log \gamma_{\text{bivalent ion}} = 2 \log \gamma_{\text{univalent ion}}$, or $(\gamma\ \text{II}) = (\gamma\ \text{I})^2$ hence the expression (3)

$$\log K_1 = \log m_{AgCl_3^=} + \log \gamma_{AgCl_3^=} - 2 \log m_{Cl^-} - 2 \log \gamma_{Cl^-} \qquad (6)$$

reduces to

$$\log K_1 = \log m_{AgCl_3^-} - 2 \log m_{Cl^-} \qquad (7)$$

This of course is only true at very low molalities and can be expected to deviate at high concentrations.

The silver chloride may be considered to be all in the form of the silver chloride complex at concentrations above 0.1 molal potassium chloride (*4*). There may be a very small amount of undissociated silver chloride present but its amount is probably so low that it is only a very small fraction of the total concentration of dissolved silver chloride. This does not entirely eliminate the possibility of the presence of $AgCl_2^-$ which probably is present in small amounts, especially

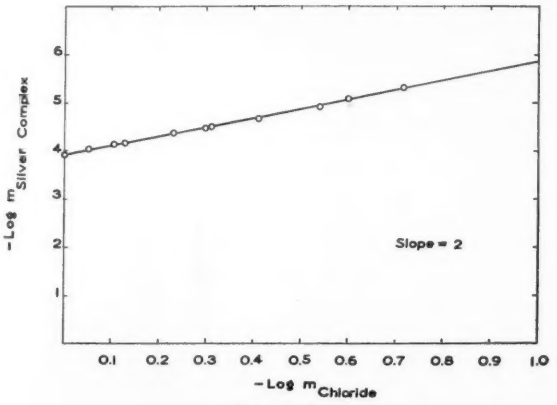

FIGURE 4

EQUILIBRIUM CONSTANT VALUES FOR COMMON COMPOUNDS IN GROUP I AT 25°

Equilibrium	*Reference*	*Constant*
$PbCl_{2\,(solid)} = Pb^{++} + 2Cl^-$	(*22*)	$K = 1.7 \times 10^{-5}$
$AgCl_{(solid)} = Ag^+ + Cl^-$	(*22*)	$K = 1.7 \times 10^{-10}$
$Hg_2Cl_{2\,(solid)} = Hg_2^{++} + 2Cl^-$	(*22*)	$K = 1.1 \times 10^{-18}$
$PbSO_{4\,(solid)} = Pb^{++} + SO_4^=$	(*22*)	$K = 1.8 \times 10^{-8}$
$Pb(OH)_{2\,(solid)} = Pb^{++} + 2OH^-$	(*22*)	$K = 2.8 \times 10^{-16}$
$AgOH_{(solid)} = Ag^+ + OH^-$	(*22*)	$K = 2.0 \times 10^{-8}$
$Pb(OH)_{2\,(solid)} = H^+ + HPbO_2^-$	(*9*)	$K = 7.3 \times 10^{-16}$
$AgOH_{(solid)} = H^+ + AgO^-$	(*10*)	$K = 7.9 \times 10^{-13}$
$PbCl_{2\,(solid)} + Cl^- = PbCl_3^-$	(*This paper*)	$K = 4 \times 10^{-4}$
$AgCl_{(solid)} + 2Cl^- = AgCl_3^=$	(*This paper*)	$K = 1.3 \times 10^{-4}$
$Hg_2Cl_{2\,(solid)} + Cl^- = Hg_2Cl_3^-$	(*This paper*)	$K = 7.9 \times 10^{-5}$
$Ag(NH_3)_2^+ = Ag^+ + 2NH_3$	(*15*)	$K = 6.05 \times 10^{-8}$

in the solutions of $m_{Cl^-} = 0.1$ approximately. Nor does it exclude the possibility of the $AgCl_4^=$ ion at high concentrations. However, these data indicate that the chief ion present is $AgCl_3^=$ in this concentration range. This is in disagreement with another conclusion (7) in which it is claimed (with little evidence) that the chief ion formed is $AgCl_5^=$.

(*b*) *Complex Ion of Hg_2Cl_2.* By the aid of the data on the solubility of mercurous chloride in sodium chloride and in calcium chloride solutions (Table 2) which are in agreement to 2 molal (chloride concentration), the character of the complex ion is found to be $Hg_2Cl_3^-$, for the slope (Figure 5) of the plot of

TABLE 1

SOLUBILITY* OF AgCl IN SOLUTIONS OF KCl AT 25°C.

Moles KCl per 1000 g. H_2O	$-\log m_{KCl}$	*Moles AgCl per 1000 g. H_2O*	$-\log m_{AgCl}$
0.000[p]	....	14.2×10^{-6}	....
0.00100[p]	....	1.12×10^{-6}	....
0.01002[p]	....	8.5×10^{-7}	....
0.1002[p]	....	1.81×10^{-6}	....
0.192	0.7167	4.72×10^{-6}	5.3261
0.252[p]	0.5986	8.33×10^{-6}	5.0794
0.290	0.5376	11.9×10^{-6}	4.9245
0.388	0.4112	20.9×10^{-6}	4.6739
0.486	0.3134	31.4×10^{-6}	4.5031
0.502[p]	0.2993	34.2×10^{-6}	4.4660
0.585	0.2328	43.4×10^{-6}	4.3625
0.741	0.1302	70.5×10^{-6}	4.1518
0.783	0.1062	72×10^{-6}	4.1427
0.882	0.0545	93×10^{-6}	4.0315
0.987[p]	0.0057	119×10^{-6}	3.9245

[p] = Data given by Pinkus (*4*).

* The unmarked data were obtained by Mr. Sidney Miller by the general method described by Hogge and Garrett for the solubility of thallous chloride in solutions of potassium chloride (*see J. Am. Chem. Soc.*, **63**, 1089 (1941)). The silver chloride was precipitated, washed, sealed with standard solutions of potassium chloride, equilibrated for a 10-day period, and filtered in the absence of light. The most of the Cl^- was removed by slow evaporation with sulfuric acid. The solutions were then made up to 100 ml. with distilled water and titrated by the standard potentiometric procedure using a silver electrode and potassium iodide; the titrations were carried out in the presence of nitric acid to eliminate lack of asymmetry in the titration curve (*see* KOLTHOFF AND LINGANE, *J. Am. Chem. Soc.*, **58**, 2457 (1936)).

log m_{Cl^-} log vs. $m^{n}_{Hg_2Cl_{2+n}}$ is unity (see a description of this method given under (*a*) above). The reaction is

$$Hg_2Cl_{2(s)} + Cl^- = Hg_2Cl_3^- \qquad K_8 = 7.9 \times 10^{-5} \qquad (8)$$

(*c*) *Complex Ions of Lead Chloride.* Here the problem is difficult and inconclusive, due to the high chloride ion concentrations at which these ions are formed. The KCl-$PbCl_2$ data cannot be used, for there is evidence of solid phase change in that system. The NaCl-$PbCl_2$ data seem best but here apparently association of sodium chloride is taking place for the values of γ_{NaCl} are increasing. Here all assumptions about the use of the limiting law are void and all calculations must employ measured values of γ. Fromherz (*8*) gives the distribution of lead in the different forms in a saturated solution of lead chloride at 25°C. as 5 per cent $PbCl_2$, 48 per cent $PbCl^+$, and 48 per cent Pb^{++}. This amount of $PbCl_2$ can also be obtained by extrapolating the graph in Figure 3 to $m_{Cl} = 0$. Using values of γ_{NaCl} and γ_{PbCl_2} in sodium chloride solutions and the value of the undissociated lead chloride given by Fromherz, a plot of log a_{Cl^-} vs. log a_{PbCl_2} can be constructed. However, the slope is approximately 0.8 from which one might conclude without too much certainty that the probable reaction is

$$PbCl_{2(s)} + Cl^- = PbCl_3^-, \text{ and } K_9 \cong 4 \times 10^{-4} \qquad (9)$$

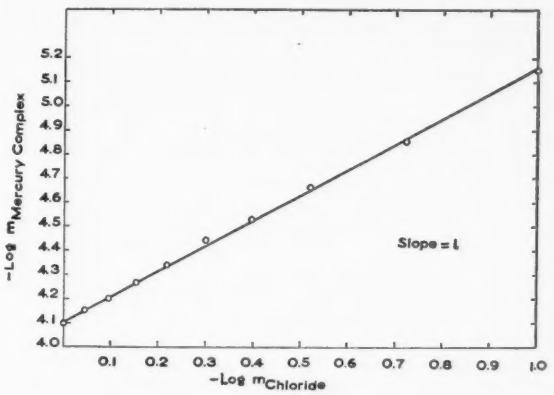

FIGURE 5

in the range of $m_{NaCl} = 2$ to 5. Certainly above that value of m_{NaCl} it is useless to attempt any quantitative treatment of the data.

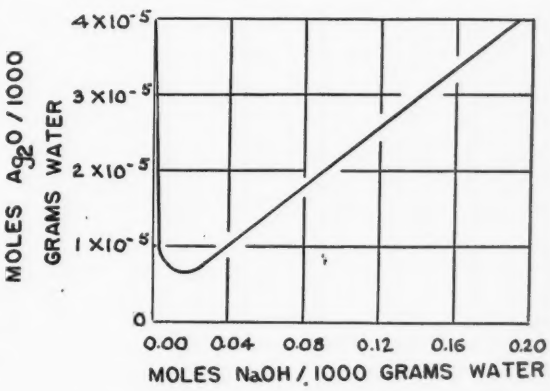

FIGURE 6

AMPHOTERIC NATURE OF LEAD OXIDE AND SILVER OXIDE

The solubility of lead sulfate and lead chromate in sodium hydroxide is explained in terms of the amphoteric nature of lead hydroxide. It is of interest to note the amphoteric nature of both lead oxide and silver oxide, which is shown from recent data (*9*, *10*) shown graphically in Figures 6 and 7. In both cases it is evident that water solutions of these oxides are predominantly basic, for the addition of hydroxyl

ions has a pronounced common ion effect. However, the increase in solubility at higher concentrations of alkali is attributed to the reactions:

$$OH^- + PbO_{(solid)} = HPbO_2^- \quad (10)$$

$$2OH^- + Ag_2O_{(solid)} = 2AgO^- + H_2O \quad (11)$$

in which the lead and silver are present in the anions. That these are the two reactions solely responsible for

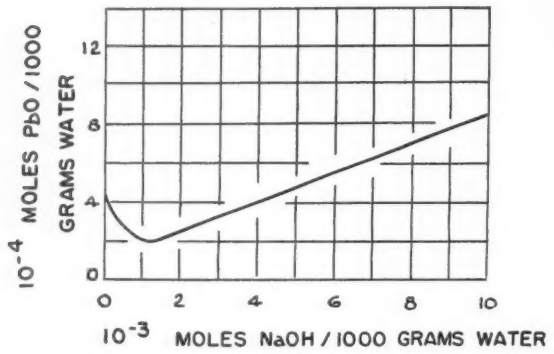

FIGURE 7

the increase in the solubility of these oxides in high concentrations of alkali can be shown by the constancy of the values of the equilibrium constants K_{10} and K_{11} (see values of K_{10} and K_{11} in data tables (*9* and *10*), and note linearity of Figures 6 and 7) over a wide range of alkali concentration. This identifies the silver and lead ion species present in high alkali concentration (to 1 molal) as the AgO^- and $HPbO_2^-$ ions.

THE SOLUBILITY OF LEAD SULFATE IN AMMONIUM ACETATE

This is explained by inference (nevertheless quite reliably) by the recent work (*11*) on the solubility of

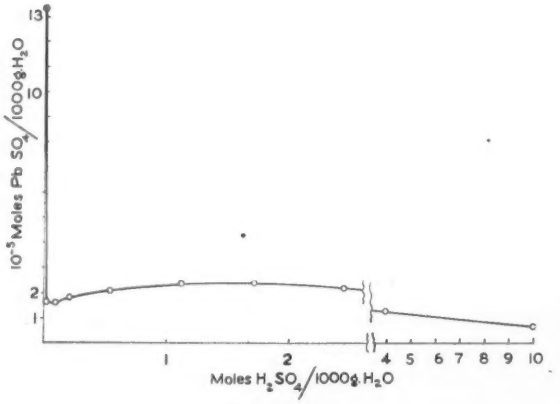

FIGURE 8

lead iodate in ammonium acetate solutions. By means of those data the authors studied the equilibrium:

$$Pb(IO_3)_{2(solid)} + nC_2H_3O_2^- = PbC_2H_3O_{2n}^{+2-n} + 2IO_3^- \quad (12)$$

It was shown that $n = 1$, by the method described under (*a*) above. This verifies the reaction and identifies the lead ion species present as $PbC_2H_3O_2^+$. The solubility is attributed, therefore, to the formation of the slightly dissociated $PbC_2H_3O_2^+$ ion. It is reasonable to assume that the solubility of lead sulfate is due to the same reaction:

$$PbSO_{(solid)} + C_2H_3O_2^- = PbC_2H_3O_2^+ + SO_4^- \quad (13)$$

In fact, this conclusion may be verified from available data (*12*) on the solubility of lead sulfate in ammonium acetate, but the conclusions are more reliable from the lead iodate data which were obtained at constant ionic strength, maintaining a constant activity coefficient ratio.

THE SOLUBILITY OF LEAD SULFATE IN SULFURIC ACID

Figure 8 shows the data on this problem (*13*) recalculated to a molality basis. These data indicate: (*a*) the optimum concentration of sulfuric acid to give the lowest solubility of lead sulfate is at approximately 0.03 molal, and (*b*) the increase in solubility of lead sulfate is only very slight beyond this minimum on further addition of sulfuric acid. This shows that the tendency of the molecule $PbSO_4$ to form complex ions is probably very slight.

HIGHER VALENCIES OF SILVER

Recent work (*14a*) indicates that silver may be oxidized to the valence state of 2 by means of ozone in nitric acid solution. This divalent state is shown conclusively by the magnetic susceptibility of solutions of the oxidized silver. The existence of trivalent silver in Ag_2O_3 in alkaline solution has been reported earlier (*14b* and *c*).

THE SILVER-AMMONIA COMPLEX

This has been reinvestigated (*15*) by measuring the solubility of silver chloride in solutions of different concentrations of ammonia. It was found that the value for the constant $K_{14} = 7.1 \times 10^{-2}$ for the reaction:

$$AgCl_{(solid)} + 2NH_3 = Ag(NH_3)_2^+ + Cl^- \quad (14)$$

and remains constant over a wide range of NH_3 concentration; hence the ion species must be $Ag(NH_3)_2^+$. The value of $K_{15} = 6.05 \times 10^{-8}$, $\pm 0.07 \times 10^{-8}$ was found for the reaction:

$$Ag(NH_3)_2^+ = Ag^+ + 2NH_3 \quad (15)$$

Evidence is presented (*16*) for the formation of $Ag(NH_3)^+$ in dilute solutions; this would not be appreciable in concentrated solution.

Caution: It should be noted that great care must be observed in handling silver-ammonia solutions. In the first place, the solution resulting from the reaction of silver nitrate with ammonium hydroxide contains ammonium nitrate which is a dangerous compound because of its explosive character even in aqueous solutions. The values of $\Delta F_{298} = -63{,}600$ and $\Delta H = -28{,}000$ for the probable decomposition:

$$NH_4NO_{3(s)} = 2H_2O_{(g)} + {}^1/_2O_2 + N_2 \qquad (16)$$

are given by Lewis and Randall (*17*). Furthermore it is dangerous to attempt to dissolve silver oxide in ammonium hydroxide in the absence of ammonium salts because of the possible formation of fulminating silver, which is highly explosive (*18*). Students should be warned against allowing such solutions to stand for a long period of time or in evaporating them to dryness.

THE REACTION OF Hg_2Cl_2 WITH NH_4OH

Recent work (*19*) indicates that the reaction product is a mixture the composition of which depends upon the concentration of the ammonium hydroxide and the time of reaction. With dilute ammonium hydroxide mercurous oxide is first formed and can be identified; the reaction ceases with the formation of $HgNH_2Cl$ and Hg. With concentrated ammonium hydroxide the water-soluble $HgCl_2{\cdot}2NH_3$ is also formed. It is concluded that under ordinary conditions the reaction between mercurous chloride and ammonium hydroxide is best represented by the equation:

$$2Hg_2Cl_2 + 4NH_4OH = Hg + Hg_2O + HgNH_2Cl + 3NH_4Cl + 3H_2O \qquad (17)$$

No new data seem to be available for the equilibrium:

$$HgNH_2Cl + NH_4^+ + Cl^- = HgCl_2{\cdot}2NH_3 \qquad (18)$$

since the work of Franklin in 1907 (*20*). At that time he indicated the equilibrium was established at 0.49 molar ammonium chloride.

TABLE 2

A. SOLUBILITY OF Hg_2Cl_2 IN SOLUTIONS OF NaCl AND OF $CaCl_2$ AT 25°C.

Moles of NaCl per 1000 g. H_2O	*Moles of $CaCl_2$ per 1000 g. H_2O*	*Moles of Hg_2Cl_2 per 1000 g. H_2O*
0.00	0.00	5.9×10^{-6}
0.101	...	8.7×10^{-6}
...	0.374	48.5×10^{-6}
...	0.528	74.2×10^{-6}
1.02	...	88.4×10^{-6}
...	1.05	170×10^{-6}
...	1.295	259×10^{-6}
2.12	...	185×10^{-6}

B. ROUNDED VALUES OF $m_{Chloride}$ AND $m_{Hg_2Cl_2\ Complex}$ TAKEN FROM LARGE SCALE GRAPH OF THE DATA IN PART A ABOVE

Moles Added Chloride per 1000 g. H_2O	$-\log m_{Chloride}$	*Moles* of Hg_2Cl_2 Complex per 1000 g. H_2O*	$-\log m_{Hg_2Cl_2}$
0.1	1	7.1×10^{-6}	5.2147
0.2	0.6990	14×10^{-6}	4.8539
0.3	0.5229	22×10^{-6}	4.6576
0.4	0.3979	30×10^{-6}	4.5229
0.5	0.3009	38×10^{-6}	4.4202
0.6	0.2218	46×10^{-6}	4.3372
0.7	0.1549	54×10^{-6}	4.2676
0.8	0.0969	62×10^{-6}	4.2076
0.9	0.0458	70×10^{-6}	4.1549
1.0	0.0000	79×10^{-6}	4.1024

* The solubility of Hg_2Cl_2 is due to the two reactions:

(1) $Hg_2Cl_{2(s)} + H_2O \rightleftharpoons Hg_2Cl_{2(aq)}$ which is 5.9×10^{-6} moles per 1000 g. H_2O.

(2) $Hg_2Cl_{2(s)} + n\,Cl^- \rightleftharpoons Hg_2Cl_{2+n}^{-n}$ which is determined by subtracting the water solubility (5.9×10^{-6}) from the total solubility giving these values in column 3.

EFFECT OF LIGHT UPON SILVER CHLORIDE PRECIPITATE

Recent work (*21*) indicates that if hydrochloric acid is in excess when silver chloride is precipitated in the presence of sunlight, the amount of silver chloride obtained in a quantitative analysis of silver is less than the calculated amount. It is concluded that the precipitate obtained during exposure to light probably consists of a mixture of metallic silver and silver chloride.

TABLE 3

SOLUBILITY OF $PbCl_2$ IN SOLUTIONS OF NaCl AT 25°C.

Moles NaCl per 1000 g. H_2O	*Moles $PbCl_2$ per 1000 g. H_2O*	*Moles NaCl per 1000 g. H_2O*	*Moles $PbCl_2$ per 1000 g. H_2O*
0.00	39.05×10^{-3}	1.71	7.72×10^{-3}
0.0856	26.03×10^{-3}	2.45	12.7×10^{-3}
0.151	15.17×10^{-3}	2.74	14.9×10^{-3}
0.342	7.54×10^{-3}	3.42	21.6×10^{-3}
0.684	6.21×10^{-3}	4.11	31.9×10^{-3}
0.856	6.03×10^{-3}	4.79	49.7×10^{-3}
1.025	6.32×10^{-3}	5.13	62.4×10^{-3}
1.367	6.78×10^{-3}	5.82	95.3×10^{-3}

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How Annabella Learned the Facts about pH

CONSTANCE S. SAMMIS *San Gabriel, California*

"YOU KNOW the old definition about gasoline," remarked the bald but learned Professor, leaning back in his old leather armchair, and wiping his spectacles. "Gasoline is the stuff which if you ain't got any in your car it doesn't run as well as if!

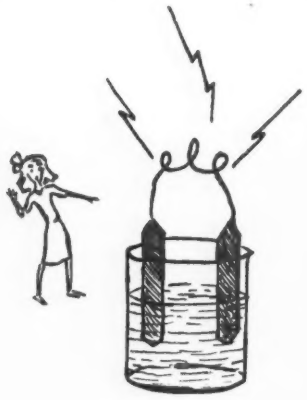
"Well, that's good as far as it goes, but it doesn't go *far* enough. It's the same idea if you say a pH meter measures pH. That is true, but it doesn't go far enough. What really happens when you use a pH meter?"

"What does?" wondered Annabella, raising one eyebrow.

"You just sit down there on that footstool, and *listen* and you'll know soon enough!" answered the Professor. So while the golden-haired Annabella settled herself to listen, the Professor continued:

AN EXPERIMENT

"Now, when you put two pieces of metal in a glass of vinegar (or a glass of soda, or a glass of most anything else that's liquid), and connect their outside ends with an electric circuit, such as a battery or dynamo—presto! you get some actual electricity flowing through the vinegar. Wonderful thing, isn't it, Annabella? And because they have something to do with electricity, we call those pieces of metal sticking in the vinegar: *electrodes*.



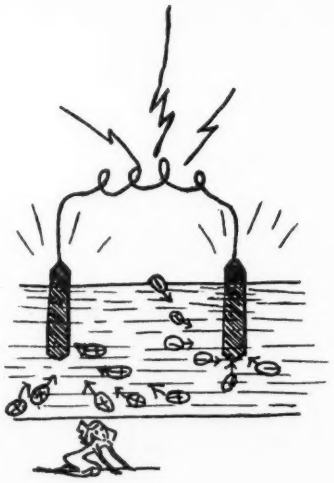
"Then, if you'll stop to think about it, you'll see that electricity couldn't flow, and keep on flowing, unless something is happening *inside* the vinegar to carry the electricity through."

"What is happening?" cried Annabella.

"What's happening is this," answered the Professor patiently, "the acetic acid in the vinegar (that's the stuff that makes vinegar taste sour) is made up of at least two kinds of tiny particles which are called '*ions*': hydrogen ions (H^+) and acetate ions (Ac^-).

"Just because it's the nature of the little beasties, the tiny hydrogen ions always have a positive electrical charge on them, and so, to balance the hydrogen ions, the acetate ions have a negative electrical charge on *them*. Everything is balanced up perfectly.

"But—when the two metal electrodes, connected to the battery, are thrust into the peace and quiet of the vinegar, things begin to happen! The hydrogen ions (H^+) begin to move off toward the electrode on the negative side of the circuit, called the cathode, and



the acetate ions (Ac^-), not to be outdone, gallop off in the other direction toward the positive electrode, the anode. How's that for breaking up a happy home?"

"Terrible," murmured Annabella, shaking her curly golden head.

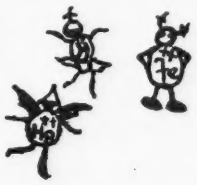
"So now you see," beamed the Professor, "it's the traveling of these little electrically charged particles that carried the stream of electricity through the vinegar, making a complete circuit from one electrode to the other. And while we're on the subject, it's easy to see that the more hydrogen and acetate ions in the glass of vinegar (the more acid the vinegar is), the stronger the *force* of the electricity that will flow through the wire.

WHAT DOES A pH METER DO?

"But don't feel badly, Annabella. A pH meter is

really a very harmless kind of instrument. It breaks up no happy homes at all. You'll see what I mean when I explain about *hydrogen electrodes*. Now—hold on tight—here we go. It hasn't been hard yet, has it? Just keep on thinking straight and this will be easy too.

"We are going to put a metal electrode like, for instance, *iron* in a liquid containing some *iron ions* (there are almost as many varieties of our new acquaintances, 'ions,' as there are people. In fact, every element and a lot of combinations of elements can exist in the form of ions). When the iron electrode is thrust into the midst of the iron ions, instantly a certain tension or '*potential difference*' springs up between the solid iron electrode and the tiny iron ions in the liquid. Reminds me of a brother and sister I had in the same class once. There was quite a potential difference between them, believe me! A sad family situation.... But, as I was saying, a tension instantly springs up between the solid iron and the iron ions, and the force of this tension depends upon the concentration of the little ions in the liquid. That's easy enough to see, I'm sure. For any certain concentration of iron ions, this tension or potential difference, as it's more often called, between the iron and its ions is *always the same* for any given temperature.



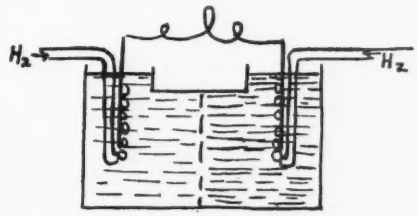
"And that's not all. If some other metal than iron is used—say, zinc or copper—the same thing happens, except that the tension set up between the metal and *its* ions will not be the same as in the case of iron, but will be greater or less depending upon the kind of metal. In other words, this tension, or potential difference, depends either upon the concentration of ions in the solution or upon the kind of metal in the electrode.

"Next, let's go back to our acetic acid in vinegar again. Now, look carefully! Suppose our two electrodes which we dip into the vinegar are not the same, but of different metals. Then the electrical tension at one electrode will be greater than that at the other. It is true that these metal electrodes won't be immersed in a solution of their own ions as was the electrode in the situation we've just been talking about, but nevertheless, a certain tension will be set up between the solid metal electrodes and the acetic acid ions in the vinegar. And because these two electrodes are of different metals, the tension between one of them and the acetic acid ions will differ from that between the other electrode and the acetic acid ions. The electrical tension at one electrode will be greater than at the other—one will be positive and the other negative, that is to say. They will generate their own electricity, really, and we won't need to connect them with a battery, but only with a wire between them, in order to get a continuous flow of current. Remember, then, we can have a potential difference between an electrode and ions, and a potential difference between two electrodes. You'll see how important it is to remember this when we talk about hydrogen electrodes.

"But a metal electrode like iron or copper is not the only kind of electrode we can have. If you bubble hydrogen gas (you know—the gas they often fill balloons with) up through a liquid and around a platinum wire that's been coated with platinum black to make the hydrogen stick to it, you'll have a *hydrogen electrode*. The same story holds true between the hydrogen gas (H_2) and the hydrogen ions (H^+) as between the iron (Fe^0) and the iron ions (ferrous) (Fe^{++}): a tension is immediately set up between them—a characteristic tension which is always and forever the same if you have the same concentration of hydrogen ions and plenty of hydrogen gas (under the same pressure). That's not hard to see, is it Annabella?"

"Oh no, no, Professor. Go on," urged Annabella, as she listened intently. She was *so* glad to be learning about the family life of her new acquaintances!

"Now if we could measure this potential difference that springs up between the electrode and its ions," continued the Professor, "that would give us a way to



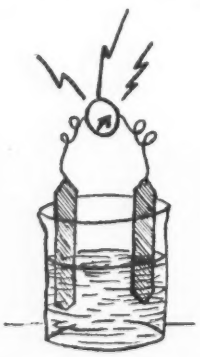
discover what the concentration of the hydrogen ions is in the liquid, *since the potential difference depends on the concentration of these little fellows called ions*. But we've found no earthly way as yet to measure this potential difference between just one electrode and its ions. So I'll tell you what we do to find out this concentration of ions. We put *two* hydrogen electrodes into solutions of their ions, with the concentration of ions surrounding one of the electrodes greater than that around the other electrode. And then we promptly connect the outside of the two electrodes with a wire, and presto, what do we have? Well, the potential differences of the two electrodes are not the same, because the ion concentrations are not the same, and therefore a current will immediately flow between the two electrodes through the wire, just as it does when we insert two electrodes of a different metal in some vinegar. That's because nature always seems to be trying to even things up—to reach an equilibrium as the scientists would say. Just like water trying to find its own level. So, since one of the electrodes is

more positive than the other, an electrical current will flow from the more negative electrode to the one that's more positive.

"That's about the whole story in a nutshell. The force of the current flowing between the two electrodes is going to equal the sum of the potential differences of each electrode. Now here's the easy part: the scientists have arbitrarily decided to call the potential difference of a hydrogen electrode dipping into a hydrogen-ion concentration of one normal (*i. e.*, one gram of hydrogen ion in 1000 cc. of solution) *zero*.[1] So all we have to do is to measure the force of the current flowing in a wire between this electrode and any other hydrogen electrode, and immediately we have the potential difference of this other electrode. That's simple enough, isn't it, Annabella? And after that, it's only one easy step to figure the concentration of ions surrounding this other electrode from a simple equation that translates potential force into ion concentration."

Here the Professor paused and looked down at the golden-haired Annabella sitting at his feet. "Do you follow me, Annabella, or have I left you stranded somewhere along the way?"

"Yes, Yes, I think I've been following you all the way," said Annabella thoughtfully. "I'm glad you went slowly though. Because it's such a new idea to me, you see. But I think I see the idea of 'potential difference.' It's when there's more of an electrical charge on one thing than on another, isn't it? And the current flowing in the wire is sort of trying to even things up between the two electrodes, isn't that it?"



"Yes, that's pretty good, Annabella. And remember that the only thing that causes the difference in charge on, say, two hydrogen electrodes (which are both the same) is the difference in hydrogen-ion concentration at the two electrodes. And believe it or not, all a pH meter does is to measure the force of this current that tends to flow through a wire connecting two electrodes each of which is thrust into a solution of its ions. All it amounts to, in effect, is a voltmeter to measure this current's force."

"Is *that* all?" exclaimed Annabella. "But, oh, I almost forgot—you said you'd explain why a pH meter never breaks up the happy homes of the little ions, Professor, remember? Will you tell me why the families of little ions don't separate and go rushing off toward these electrodes, like we talked about with the vinegar?"

[1] "The potential difference between a hydrogen electrode under one atmosphere pressure of hydrogen and a hypothetical solution normal with respect to the hydrogen ion shall be considered zero at all temperatures."—W. M. CLARK in "The determination of hydrogen ions."

"Why yes, certainly, Annabella. It's really quite simple," answered the old Professor kindly. And pausing only long enough to take out an enormous white handkerchief from his coat pocket and to blow his nose loudly, he serenely continued. "You see, if we allowed the current actually to flow through the wire, and the ions to carry this current through the solution by traveling each to its respective electrode, the solution itself would be changed by losing these ions—that is, the concentration of these ions would be constantly changing. What's the use of measuring ion concentrations if they change as fast as you measure them? But, luckily, some very smart scientific men have thought of a way to overcome this difficulty."

"How in the world *could* they?" wondered Annabella.

"It's really a pretty smart idea," continued the Professor, "for all it's so simple. And this is it: instead of letting the current flow in the wire, they promptly oppose this current by another current from a battery which is made to have the *same force* as that flowing between the electrodes—and consequently the two currents bump headlong into each other and quite cancel each other out! All they do is run in a battery current of enough force to cause the meter to show absolutely nothing going through the wire, and that means that the two opposing forces are exactly equal, you see. Then they measure how much the force of the battery current had to be to cause this deadlock—and there you have it—*that's* the force of the current that tries to flow (but can't) between the two electrodes. And thus the concentation of ions in the solution is saved from change. This clever arrangement is known as the *potentiometer*."

"Well, what do you know!" exclaimed Annabella. So that's how they find out the force of the current without breaking up any happy homes!"

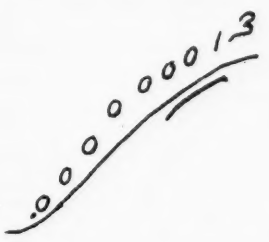
"A minute ago," the Professor added, "we found out that the force of this current tending to flow between the electrodes will depend on the concentration of hydrogen ions in the liquid we're measuring. And so, as we discovered, measuring this force is really at the same time measuring the concentration of hydrogen ions in the liquid. We translate this electrical force into hydrogen-ion concentration by a fairly simple equation, which you don't need to bother your little head about now. But the nice thing to know is that this is all done automatically for us by a pH meter. You just read the pH of the liquid you're testing right off the dial of the pH meter."

"But what's pH?" asked Annabella. "You haven't introduced me to that."

pH

"The 'pH' of a liquid is just the way the scientists have decided to express the hydrogen-ion concentra-

tion. Goodness knows why, but they do! One reason, I suppose, is that if they actually told what the concentration of hydrogen ions was (in terms of 'mols per liter'—the ordinary way scientists would express it) they'd have to put down something like this: .000000013. And pH is merely a simpler way of saying what the hydrogen concentration is.[2] For instance, translating this long number into pH units, it would be just 7.8. You'll have to admit that pH is an easier number to say. And remember this, the *more* hydrogen ions (that is, the more acid the liquid), the *lower* the pH. That's confused a lot of people, I'm afraid. But if you'll keep in mind that the *higher* the hydrogen-ion concentration, the *fewer* the pH (and vice versa), you'll have it well in hand.



"Now, take a look at the dial on this pH meter here on my desk," directed the Professor to Annabella, pointing to it with his spectacles. "See, everything that shows a pH from 0 over to 7 will be *acid*. A liquid that tests *above* pH 7 will be *over on the alkaline side*. And right in the middle—at pH 7—what do you think? Why, pure water, of course! It's neutral."

"Of course!" echoed Annabella, gazing at the dial with satisfaction now that she knew what it meant.

JELLY, AND THINGS LIKE THAT

"And fundamentally, that's *all* there is to it. Now a few more or less hydrogen ions can cause an awful lot of difference when you're making something. Take jelly, for instance. If you've ever watched your mother stand over a kettle of hot fruit in the summer-time and concoct a batch of beautiful jelly, you'll know it isn't always easy to make the juice jell to just the right stiffness.

"Well, the villain in this situation is the *hydrogen* ion! At a pH of 3.2 the jelly will be just the right stiffness—firm but not too hard. But at a pH of 3.3 (a tiny bit more on the alkaline side) the jelly is tender; at pH 3.4 it's much too soft; and at pH 3.5 the juice won't jell at all! Going in the other direction, it's been found that at pH 2.8 (a little more acid) we get that 'weeping effect' that all jelly makers bemoan. If we can test for and find out the pH of the juice then we can easily control its acidity.

"Oh, if there were only a pH meter in every kitchen!" breathed Annabella. "What jelly we could turn out!"

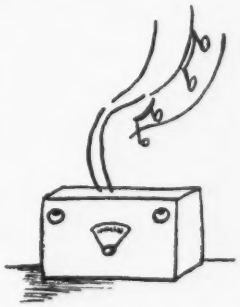
"Yes, but seriously now, my dear," said the old Professor earnestly, "you can see how important the meter would be to manufacturers who make jelly, can't you? It means they can control the pH of their product with it. They're not haunted by the fear of ruined batches of jelly being sent back by angry grocers. And this is only one example out of thousands where hydrogen control is important. They practically never make cheese in this country now without the help of a pH meter. Hydrogen-ion control is important in making sugar, rayon, paper, aluminum for airplanes, and many other things."

"BUT WHAT'S THIS TALK ABOUT A GLASS ELECTRODE?" asked Annabella timidly.

"Oh, so you're wondering about that too? Well, that's easy," replied the Professor. "But I'll have to be brief." And here he took out his heavy gold watch from his vest pocket and laid it carefully on the desk top. "My chemistry class will be waiting for me in five minutes, you know."

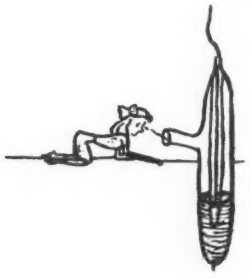
"Hydrogen electrodes are often very bothersome to use. You have to manufacture hydrogen for them, and they don't get along at all well with some liquids. They are still the standard, though, please remember. After our friends the scientists had done a lot of experimenting to find a good substitute for them, somebody thought of making a *glass electrode*. That may sound like a funny thing to use, but actually it works extremely well. It can be dipped in any kind of solution and is practically immune to all chemicals, no matter how strong. There is a certain potential difference that springs up between the glass and the hydrogen ions which is just as definite and measurable as that between hydrogen gas and hydrogen ions.

"So the glass electrode was a wonderful idea, except that the current set up when glass electrodes are thrust into solutions of hydrogen ions is very, very small. It can be measured, but it is much too small to measure easily. And *now* what they do (they're always thinking up something, aren't they?) is to *amplify* this tiny current up to where they can measure the force of it on an ordinary voltmeter. Just the way your radio am-



[2] $pH = -\log (H^+)$.

plifies weak signals coming in from the air up to a point—well—where your next door neighbor complains! Of course the pH meter amplifies just the same amount all the same time, so they can still measure the changes in voltage that different ion concentrations produce.



"The *calomel electrode* is the other electrode that is used in most pH meters along with the glass electrode now. It is made of glass too, but it has an inner electrode of mercury and some calomel and potassium chloride inside it. The mercury really acts as an electrode dipping into the mercury ions of the calomel, and the potassium chloride solution is what comes in contact with the liquid we want to measure. But all you need to remember about the calomel electrode now is that it is used as a reference electrode because its potential difference is practically the same, no matter what type of ions it is thrust into.

AND SO:

"The main thing to remember," concluded the Professor, carefully wiping his spectacles and putting them on the bridge of his nose again, "is that a pH meter measures the force of an electrical current that *tends* to flow (that *would* flow if it were not for the potentiometer) when two connected electrodes are plunged into a liquid. The force of this current will tell us how acid that liquid is. That is, it tells us what the concentration of the hydrogen ion is. And this information is very, very important to industrial users and research men. It helps them save a lot of money by turning out a more uniform product with less supervision. It takes the guesswork out of many of their operations.

"I'll bet you thought it was a lot harder than that, didn't you now, Annabella? Of course, working out the amplification circuits, the special kind of glass electrodes, and other details did require a lot of hard work on the part of the research men. But don't you worry about those things. Don't let them cloud your idea of how a pH meter works. Just stick to that one simple idea: measuring the acidity by measuring the force of the current that tends to flow between two connected electrodes in a liquid. And the acidity depends on the hydrogen-ion concentration.

"And now, my dear, you really *know* what goes on when somebody twists the knobs of a pH meter, and don't let anybody ever tell you that you don't," finished the old Professor firmly, adjusting his spectacles securely to the bridge of his nose, replacing his watch in the pocket of his ample vest, and rising slowly to his feet. "I hope my chemistry class hasn't grown tired of waiting for me."

"Oh, thank you, Professor, how can I ever *thank you*!" cried Annabella. "You don't know how much I've learned today!" And waving farewell, off she flitted—to buy a pH meter for her mother's kitchen.

LETTERS

Ounce Molecular Weight

To the Editor:

I have read with interest and approval the suggestion in the June issue of the JOURNAL OF CHEMICAL EDUCATION by Professors Audrieth and Johnstone that students in general chemistry should be given more experience in the use of the English system of weights and measures in making chemical calculations. But I do not think that it is wise to require the student to learn a new number, 359, which is the pound molecular volume in cubic feet.

A few years ago I either read in some forgotten book or article or noted by chance that the volume in cubic feet occupied at standard conditions by an ounce molecular weight of a gas is 22.4. The identity of this with the gram molecular volume in liters is of course due to the accidental near equality of the number of grams in an ounce and the number of liters in a cubic foot. In teaching freshman chemistry I have made it a habit to point out to the students that 22.4 is by chance not only the gram molecular volume in liters but also the ounce molecular volume in cubic feet, and I have given the students practice in working gas-law problems in the English system as well as the metric system. Use of the English system seems to give the students an increased feeling of the significance of the calculations that he is making.

LINUS PAULING

CALIFORNIA INSTITUTE OF TECHNOLOGY
PASADENA, CALIFORNIA

Don't Segregate!

To the Editor:

I was interested in the "Editor's Outlook" in the June number. I have had an extremely interesting experience during the last three months. After teaching forty-one years at a well-known "higher institution of learning" in Massachusetts, I was fired in 1941 for old age and senile decay, etc. After finishing up two new editions of some textbooks, I was pleased this spring to have two chances of getting back into the teaching game, which I accepted with alacrity. The head of the Science Department of a well-known preparatory school was called to Washington to work for the U. S. Navy. I was asked suddenly, without any preparation at all, to take over three classes in chemistry and one in physics. Since I have never tried to teach physics nor studied it to any extent since 1896, the proposition was probably the hardest one that has ever been put before me. . . . .

From the experiences I have had during the last few weeks, I can readily understand the letter on the examination book from the beautiful lassie whose mind was in a complete fog throughout the entire year.

The three sections that I had in chemistry were divided in the way that you suggest. The boys and girls going to college were in one division, those who never meant to go further in chemistry were in another division, and there was an intermediate class. The poorest division had a different text, with more pictures and fewer details. The aim had been, as you put it, to tell this class *about* chemistry, with the realization that they would not want to *study* chemistry to any extent. When I took the class, only about two out of the thirteen were passing the course, and they had low grades. I have never seen such an uninspired group of students in my life. From my experience with this group, I am sure that your "Utopian ideal" is a sour proposition. Trying to feed pabulum in chemistry is not a wise thing to do. As Wilhelm Ostwald pointed out about forty years ago, chemistry has always been taught to beginners from too low a level. In physics, it is assumed that the students have some grasp of mathematics, but in chemistry it is usually assumed that everything must be made easy and entertaining. Hence so many silly pictures in the texts for beginners, and so many useless films that are sent around in order to interest youth in the "wonders" of chemistry. We had a very fine projector, and I tried showing quite a number of these films based on the idea of "visual education," until one day finally a student asked: "Why do you show us those films? They teach us nothing new and you have a lot of things to explain to us."

I had only about seven weeks to work with this poor class. At the end of the first week, I found to my dismay that not a single boy or girl could follow me well in the classroom. I found that there was absolutely no chemical fact or theory that I could assume to be known. In the remaining six weeks, I gave to the class exactly the same lessons that I gave to the other sections, but taught each lesson with somewhat greater care, and made the written tests a little briefer. I stressed a logical and mathematical treatment. To my surprise, the reaction was good, and quite a few came to me and said: "I am beginning to learn something about chemistry."

In some ways the teacher of chemistry should follow the pattern of the physician. He studies the individual case of each patient and tries to prescribe suitable treatment for each. He does not have a single medicine which he prescribes to all. It is difficult in a large class in Freshman Chemistry to study each individual, but if the class is broken up into small recitation sections of not more than 25 students each, a good instructor ought to get to know each student fairly well. It is not usually possible to make every student study enough to pass the course, but this much is positively true: if the student is frank with the instructor and is willing to work hard, it is a poor instructor who cannot teach at least 90 per cent of such students. The instruction has to be more or less adapted to the needs of the class. Thus I couldn't treat beginners exactly as I would treat postgraduate students at a college or technical school. If the instructor knows his subject very well he should be able to impart his knowledge to beginners and if he can't, the presumption is that he does not know his subject very well. Of course, some scholars cannot think in words of less than three syllables, but they are not good teachers, for the most part.

My thesis, therefore, is this—if the faithful student is in a haze throughout an entire year, the fault is on both sides. The student should have consulted the instructor the moment he began to get into a fog, and the instructor should have found out about the student before many weeks had passed. . . . The secret of good teaching, in my opinion, lies in making the boys and girls work hard to learn the subject and to study intelligently. The poorer boys and girls have to be directed more, but, in the long run, the earnest students are more likely to succeed than the brilliant ones. I have worked several times with students carefully graded into sections, but, on the whole, I have had better luck with classes in which the good and the bad were well mixed. Then, often a slow student will get help from one of his brighter classmates, and sometimes the students can teach better than some of the instructors.

The thing that has impressed me most in changing from college to a preparatory school is that many of the suggestions I have read about teaching in the JOURNAL OF CHEMICAL EDUCATION were wrong.

WILLIAM T. HALL

ROCHESTER, MASSACHUSETTS

Ten thousand tons of steel went into the metal tops of the 1,000,000 jars manufactured last year to hold cold cream and other cosmetic products.

Otto Ruff

RALPH E. OESPER *University of Cincinnati, Cincinnati, Ohio*

UNPROFITABLE ventures in match manufacture and shale oil distillation by Otto Ruff's grandfather and father convinced the family that a young man in moderate circumstances had better exercise his chemical talents in a profession giving more promise of a sure livelihood. However, three years' service as apothecary apprentice was enough; Ruff felt that he was made for better things than vending packaged remedies and rolling pills. He decided to risk his future, and devote himself to chemistry.

He chose the University of Berlin because Emil Fischer was there, but, of course, Ruff had no direct contact with the great man for several years. Ruff's manipulative skill attracted attention; after only one semester he was made assistant to Piloty, under whom he completed his doctorate in 1897. Ruff passed the state examination in pharmacy in 1896, and during his studentship supported himself, for the most part, by clerking in drugstores.

The young doctor stayed on at Berlin, but now as assistant to Fischer. He changed his field of research from nitroso compounds to sugars, and achieved considerable success, particularly in his studies of the degradation of these carbohydrates by oxidation. Ruff's habilitation lecture in 1901 was entitled, "Synthesis in the Sugar Groups," and he seemed destined to become another in that great company of organic chemists that issued from Fischer's influence. This was not to be; and none other than Emil Fischer was responsible for a sudden and final change in Ruff's interests. A good organic chemist was forced into the inorganic field, and the chemical world is doubtless far richer.

From time to time Ruff took on problems brought to him by the organic industries. His work on carbazole was so good that a tempting offer to enter the full-time employ of one of the big dye corporations followed. How often have brilliant young professors been confronted with this situation? Ruff's mind was made up for him by Fischer, who offered him the headship of the inorganic division, but with the stipulation that Ruff and his assistants not only teach inorganic chemistry but do their research in that field. Fischer fully realized that the over-emphasis on the organic side was a real danger to the healthy advance of German chemistry.

Faced with the necessity of finding new fields of investigation, Ruff felt his way cautiously, getting inspiration and suggestions wherever he could. Always interested in materials rather than theories, he naturally turned to the preparative side, and soon had to his credit a series of interesting new compounds. The most important chapter in his career isused from such simple beginnings.

The determination of the titanium content of silicates was based on the volatilization of the silicon, as SiF_4, when a mixture of the oxides is treated with hydrofluoric and sulfuric acids. As titanium tetrafluoride had not been prepared, the method could not be considered certain until actual tests of the nonvolatility of this compound had been made. Ruff prepared TiF_4 by heating the tetrachloride with anhydrous hydrofluoric acid, whose preparation in those days was in itself quite a feat. The reaction went so smoothly that it was successfully tried on many other chlorides. In time, Ruff developed other methods of preparing fluorides, and eventually became the world authority on inorganic compounds of this element. Some sixty new fluorides were made in his laboratory and others were studied there. Much of this involved beautiful low-temperature technic. His biographer Hückel,[1] points out that although RuF_5 was prepared by Ruff, his efforts to make RuF_2 were not successful, which really was too bad, because the formula would have spelled his name! Ruff was an expert in demonstration experiments and his illustrated survey of fluorine chemistry before the German Chemical Society in 1936 will long be remembered as a model of this type of lecture.[2]

Like Moissan, the isolator of fluorine, Ruff also contributed much to high-temperature studies. The production of good stoneware, the manufacture of ceramic vessels for high-temperature work, synthetic gems, the various forms of carbon, electrolysis of fused salts, gas explosions in coal mines, are typical instances of his broad range of interests.

In 1904, he went to Danzig to direct work in inorganic and technological chemistry; in 1916 he was called to the Technical High School at Breslau, and from 1933 to his retirement in 1936 he was also on the faculty of the University of Breslau.

Ruff was purely a synthetic chemist; his chief aim was to make and study the physical properties of compounds. Werner's coördination theory, and similar concepts, and the application of physical chemistry appealed to him slightly if at all. He felt that theory had fallen far behind the experimental data of the inorganic field. The niceties of analytical chemistry were passed over; to him analysis was merely a means to an end. He was a "lone wolf," and founded no school. Honors came to him in full measure, both in Germany and abroad, but the 300 papers that came from his laboratory constitute his lasting memorial.

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[1] HÜCKEL, *Ber.*, **73A**, 125 (1940).
[2] RUFF, *Chem. Ztg.*, **60**, 873 (1936).

Out of the Editor's Basket

AMONG THE many articles which have been written recently on chemical warfare agents and decontamination we would call attention to one in particular. Entitled "Chemical Warfare (II). Protection. First Aid. Decontamination," it appeared in the June number of *Science Studies* from St. Bonaventure College. Its excellence is in its arrangement, rather than in any novel methods or information which it contains.

● There also came to us, from the Office of Civilian Defense, a mimeographed outline of a special civilian defense course on "Incendiary Materials" by Major W. E. Caldwell. The course originated at the War Department Civilian Protection School of the University of Maryland. A number of excellent demonstration experiments are described. For copies, send to: Protection Branch, Inspection Section, Office of Civilian Defense, Washington.

● From the University of Wichita, Kansas, comes Bulletin No. 11, entitled "Bibliography of Some Achievements in Chemistry, 1930–1941," by Lloyd McKinley, Head of the Department of Chemistry. It should find a ready use by teachers and students, for a large number of original references are grouped under 48 topical headings, which cover principally the fields of applied chemistry, emphasizing the theme: "Better Things for Better Living through Chemistry." Three sections of theoretical interest are included, however: "Absolute Zero (low temperature research)," "Atoms, Elements, Isotopes, and Radioactivity," and "Chemical Nomenclature." We can imagine a number of ways in which it would be handy for class use.

● Attracting nation-wide attendance, the second annual workshops on junior college terminal education were held this summer at the University of California, the University of Chicago, and Harvard University. The third and last of this series of workshops will be held in the summer of 1943.

Designed to "teach the teachers" so that they may adapt the increasingly popular two-year terminal courses most effectively to the needs of junior college students, the workshops have become especially valuable to the instructors during wartime, faced as they are with the educational race-against-time. The workshops have been sponsored by the American Association of Junior Colleges through its Commission on Junior College Terminal Education and have been operating under a grant from the General Education Board.

● One angle of the current drive against waste is the organized effort now being made by the War Production Board to encourage the process of solvent recovery. It is estimated that over one billion pounds of solvents are now recovered annually. Close to double this amount could be recovered, it is said, if coöperation of all producers was obtained. The variety of solvents involved includes alcohols, chlorinated compounds, esters, ethers, hydrocarbons, and ketones. These are used in large quantities by dry cleaners, rotogravure printers, metal degreasers, wire insulators, and solvent extractors—to mention but a few industries using solvents. They enter into the manufacture of smokeless powder, linoleum, "glassine paper," artificial leather, impregnated fabrics, awnings, rayon and other fibers, plastics, transparent wrappings, cellulose nitrate and acetate, and rubber.

While the most important consideration in reclaiming solvents is that of making available essential chemicals for war production, there is also a substantial saving in raw material costs involved. A solvent that costs twenty cents a gallon, for example, may be reclaimed and used again at a fraction of the original cost.

● Chemists should be especially interested in the observance of Fire Prevention Week, October 4 to 10, since combustion is one of the most important chemical phenomena—for better or for worse.

● Ordinarily a gas is not thought of as an alloy. However, it is reported that nitrogen is being used in Germany to replace nickel in high-alloy special steels.

● Lightning, which sometimes causes damage, does far more good than harm. In its passage through the air, oxygen and nitrogen are combined by the electric arc action to form oxides which unite with the raindrops to form nitric acid. In this manner the soil is enriched to the extent of 100 million tons of nitric acid annually—more than is produced by the combined output of all the world's fertilizer plants.

● After many months of experimentation the American Can Company has developed, it was recently announced, a revolutionary method for the making of cans with fiber bodies on machines used for the manufacture of metal containers. This new method, which American Can will probably make available to the entire industry as soon as the new method has been thoroughly tried and perfected through actual production, is considered the most important development within the can manufacturing industry within the past decade. The new process brings the first ray of hope to the vast number of American manufacturers of dry products whose merchandise was packed in cans, and whose business is threatened with dislocation by the various Government orders restricting the use of tin and metal containers. The greatest merit, perhaps, of the innovation is that no new machinery is required.

Report

of the

NEW ENGLAND ASSOCIATION
—of—
CHEMISTRY TEACHERS

Report of the Committee of the Fourth Summer Conference

THE FOURTH summer conference of the NEACT was held from August 11 to 14 on the campus of the University of New Hampshire, at Durham, New Hampshire. The university very generously made available all of its excellent facilities and coöperated in every way to make the conference a success. The local arrangements were very ably handled by Professor Edward R. Atkinson. His efforts, together with those of the other members of the staff of the Department of Chemistry, made the conference very enjoyable. The close collaboration of the News Bureau and the Photo-Visual Service of the University of New Hampshire added greatly to the enjoyment of the 163 conferees and guests who were present.

The program was planned, as at previous summer conferences, to afford a period of professional stimulation relating to some of the more recent advances in chemistry and allied sciences. The committee feels that the success of a conference such as ours is measured by the participation of the conferees in all of the discussions, and at this conference there was excellent response. The exchange of ideas between teachers of both the secondary schools and colleges serves, as no other means can, to aid in the improvement of chemistry teaching everywhere.

The twenty-three speakers on the program leave the Association deeply indebted to them; their contributions have made this conference a worthy example for educational conferences. The symposium was very well received, and the honest agreement and disagreement on theories aided in clarifying a difficult topic in the minds of those present. The round tables were also successful in serving to initiate ideas of collaboration which may well lead to furthering the teaching of chemistry.

The conferees and guests enjoyed a social hour at Scott Hall after the first evening program. Tea was served by the wives of the chemistry staff of the university on Wednesday and Thursday afternoons in the living room of the dormitory. Many conferees and guests enjoyed a boat trip on the bay Thursday afternoon. Due to the thinness of tires and the rationing of gasoline, not many got out to the neighboring golf courses, but several cooled off in the university outdoor swimming pool. Trips to industrial plants were not emphasized, but there were four available for those interested.

Thirteen exhibitors from educational and commercial organizations added to the value of the conference, and five industrial concerns lent films for use at the evening meetings. The staff of the Department of Chemistry of the University of New Hampshire prepared nine exhibits showing new developments in apparatus and technics for teaching purposes as well as for research.

I should like to take this opportunity to thank all members of the committee and the many other willing workers who coöperated so generously for their part in making the Fourth Summer Conference such a noteworthy activity of the Association.

—MILLARD W. BOSWORTH, *Chairman*

Note: The program for the Fourth Summer Conference, listing the principal speakers, may be found in the JOURNAL OF CHEMICAL EDUCATION for July, page 312, or in the REPORT for May, page 211.

OFFICIAL BUSINESS

THE ANNUAL meeting, which was adjourned *sine die* May 16, 1942, was continued on Friday, August 14, at the University of New Hampshire. The meeting was called to order at noon by President Theodore C. Sargent. Miss Sue C. Hamilton, treasurer, was appointed secretary, *pro tem.*

A letter from Miss Eva M. Ruggli was read in which she requested the Association to accept her resignation as Secretary. President Sargent announced that the Executive Committee had so voted, with greatest regret.

The motion of the Executive Committee to adopt the new version of the Constitution (REPORT of the NEACT, **43,** 214 (May, 1942)) was read by the secretary and was passed unanimously. The Executive Committee announced that the By-Laws had also been revised to bring them into conformity with the revised constitution. Having been authorized at the 216th meeting to complete arrangements with the JOURNAL OF CHEMICAL EDUCATION for the printing of the REPORT, the Executive Committee announced that a contract had been drawn up and executed by the proper officials of the JOURNAL and of the Association with the following provisions:

1. The JOURNAL will provide space to the extent of not less than two pages per issue for the REPORT of the NEACT.

2. The Association will purchase copies of the October, November, and December issues of the JOURNAL, and starting with the January issue each member of the Association will receive a year's subscription to the JOURNAL. Members who are already subscribers will be refunded the balance of their present subscription.

3. The JOURNAL will provide the Association with reprints of the pages reserved for the REPORT and with two complete volumes of the JOURNAL for its archives. The JOURNAL will bind 25 complete sets of the reprints, retaining the sequence of volume numbers of the REPORT.

4. The Association will nominate an Associate Editor for the REPORT pages, who shall be acceptable to the Editor of the JOURNAL. Material furnished by the Association will be subject to the JOURNAL's usual editorial procedure.

5. The agreement will remain in force until revoked by either party on one year's written notice.

President Sargent announced that the Executive Committee recommended the nomination of Professor Laurence S. Foster, Brown University, as Associate Editor of the JOURNAL OF CHEMICAL EDUCATION, in charge of the NEACT pages. The members present concurred unanimously.

NEW MEMBERS

The Membership Committee reported the following new members had been elected at the Fourth Summer Conference:

Dorothy Moody Brown, Science Instructor, St. Luke's Hospital, New York.

Frederic B. Dutton, Associate Professor of Chemistry, Baldwin-Wallace College, Berea, Ohio.

Elizabeth Gillespie, Head of Science Department, Maury High School, Norfolk, Virginia.

Reverend Leo J. Guay, S.J., Department of Chemistry, Holy Cross College, Worcester, Massachusetts

John A. Hangen, Teacher of Biology and Chemistry, Farmington High School, Unionville, Connecticut.

William B. Hounsell, Headmaster of Colebrook Academy, Colebrook, New Hampshire.

William S. Huber, Chemistry Instructor, Rhode Island School of Design, Providence, Rhode Island.

Eleanor Jones, Lachute High School, Lachute, Quebec.

Guita Marble, Instructor in Chemistry, Women's College, University of North Carolina, Greensboro, North Carolina.

Reverend Joseph A. Martus, S.J., Cranwell Preparatory School, Lenox, Massachusetts.

Raymond E. Neal, Assistant Professor of Chemistry, Simmons College, Boston, Massachusetts.

Sister Marie Jeanne-of-the-Visitation, Science Teacher, Notre Dame High School, St. Regis Academy, Berlin, New Hampshire.

Sister Marie Ste. Croix-de-Jesus, Science Teacher, Presentation Academy, Hudson, New Hampshire.

Sister Marie St. Theophile, Professor of Chemistry, Rivier College, Nashua, New Hampshire.

Helen A. Thayer, Teacher of Chemistry and Science, High School, Watertown, Massachusetts.

Seth G. Twitchell, Head of Science Department, High School, Concord, New Hampshire.

Elizabeth Whittemore, Head of Science Department, Concord Academy, Concord, Massachusetts.

Katherine K. Wimp, Academic Head, Wykeham Rise, Washington, Connecticut.

FIFTH SUMMER CONFERENCE

The Executive Committee announced that the Association would proceed with plans to hold a Fifth Summer Conference, subject to review at the first meeting of the Committee in the fall. It was voted to accept the invitation of Massachusetts State College, with the reservation that it might be necessary to change plans, due to conditions imposed by the war.

OFFICERS OF THE NEW ENGLAND ASSOCIATION AND MEMBERS OF THE FOURTH SUMMER CONFERENCE COMMITTEE

Seated: Ethyl Weymouth, Clara M. McDonald, Sue C. Hamilton, Dorothy W. Gifford.

Standing: Benjamin R. Graves, S. Walter Hoyt, Lester F. Weeks, Theodore C. Sargent, George B. Savage, John Hogg, Millard W. Bosworth, Eldin V. Lynn, Elbert C. Weaver, Edward R. Atkinson, Laurence S. Foster.

1942–1943 SCHEDULE

Unless the Association is forced to alter its program, there will be six meetings held during the year 1942–43. The dates have been chosen so that there are no conflicts with the meetings of the Eastern Association of Physics Teachers and of the New England Biology Teachers Association. The December 5 meeting, as usual, will be held jointly with these other groups.

217th Meeting, October 10, 1942, Technical High School, Springfield, Massachusetts.

218th Meeting, December 5, 1942, Simmons College, Boston, Massachusetts. (Joint meeting).

219th Meeting, February 6, 1943, Southern Division.

220th Meeting, March 6, 1943, Western Division.

221st Meeting, April 3, 1943, Northern Division.

222nd Meeting (Annual Meeting) May 15, 1943, Central Division.

SECRETARY

On the final day of the Conference, President Sargent announced that the Association had been fortunate in being able to secure the services of Professor Raymond E. Neal, of Simmons College, for the important post of Secretary of the NEACT, left vacant by the resignation of Miss Ruggli. Professor Neal was born in Lynn, Massachusetts, and attended the Lynn English High School, and was graduated from Harvard University with the class of 1919. During the last war he saw overseas service with the 301st Field Signal Battalion. He has been on the staff at Simmons College since 1920.

RECENT BOOKS

ORGANIC REAGENTS IN INORGANIC ANALYSIS. *Paul von Stein*, Director of Analytical Developments, Cadmium Residue and Pigment Department, Harshaw Chemical Company, Elyria, Ohio. Chemical Publishing Co., Inc., Brooklyn, N. Y., 1942. viii + 242 pp. 13.5 × 21.5 cm. $4.50.

This book is a compendium of organic compounds which yield indicative reactions with inorganic substances. Pertinent facts are given for each determination, as well as reference to the original articles from which the material has been taken. The elements dealt with are given in alphabetical sequence, each composing a chapter (47 in all), which is subdivided into paragraphs with the organic reagent under consideration as the subtitle. The procedures for the more important methods are given in full as regards the care and technic required; methods of less importance are mentioned only briefly. The preparation of reagent solutions and a partial summary of reagents and their interferences in detecting metals are given in the Appendix. The book is well indexed.

No attempt was made to include or describe the chemical reactions, affinities, and formulas of the organic reagents. The book will be a useful adjunct to every analyst; those interested in the theoretical aspects of the subject of metallo-organic analytical methods will need to consult the treatises and original articles, many of which have appeared within the past ten or fifteen years.

JOHN H. YOE

UNIVERSITY OF VIRGINIA
CHARLOTTESVILLE, VIRGINIA

PAINTING MATERIALS. A Short Encyclopedia. *Rutherford J. Gettens*, Chemist, Department of Conservation, and Fellow for Technical Research, Fogg Museum of Art; and *George L. Stout*, Lecturer on Fine Arts and Head of the Department of Conservation, Fogg Museum of Art. D. Van Nostrand Company, Inc., New York, N. Y., 1942. vii + 333 pp. 37 figs. 17 × 24 cm. $3.25.

The authors have attempted to compile the significant information on the materials and processes of painting, including information resulting from recent studies in the fields of chemistry, microscopy, and physics. The work is arranged as an encyclopedia, in five sections—Mediums, Adhesives, and Film Substances; Pigments and Inert Materials; Solvents, Diluents, and Detergents; Supports; Tools and Equipment—is adequately cross referenced, and includes lengthy bibliographies for each section.

The data were gathered, according to the authors, ". . . for all who do work in the art—painters, teachers of painting, students, museum curators and conservators, paint chemists and analysts. There is much that will concern the museum worker and the paint analyst more than others—distinctions among chemical and physical properties, problems of conservation, and history of materials."

CHEMICAL DICTIONARY. Compiled by *F. H. Campbell*, D.Sc., F.A.C.I. Chemical Publishing Co., Inc., New York, N. Y., 1942. 85 pp. 13 × 21 cm. $2.50.

"Chemists find it necessary to refresh their memories of meanings of terms not in constant use and to keep abreast of new terms constantly being introduced. This condensed book contains the more frequently used terms which will meet the needs of those who wish to make their written expression as precise as laboratory results."

The author has included those terms which chemists use, even though not strictly chemical, and has attempted to eliminate all terms for which potential users of the book already know the meaning. Included is a useful four-page summary of inorganic and organic nomenclature.

MINERALS IN NUTRITION. *Zolton T. Wirtschafter*, M.D. Clinician in Charge, Clinic for Diabetes, Mount Sinai Hospital, Cleveland; Visiting Physician, Cleveland City Hospital; Clinical Instructor in Medicine, Western Reserve University. Reinhold Publishing Company, New York, 1942. 175 pp. 4 figs. 12 × 18.5 cm. $1.75.

This book on mineral metabolism is intended for the lay public, and is therefore written with non-technical terminology, in a clear, simple style. After introductory chapters on the functions of minerals including water, separate chapters are devoted to each of the major mineral elements found in the body. Consideration of the mineral needs of pregnancy and lactation and a summary chapter complete the book. Interesting X-ray photographs showing demineralization of the skeleton, rachitic bones, and lack of development of bones in Cretinism illustrate the chapters on calcium, phosphorus, and iodine.

The fact that each mineral is considered separately means that there is considerable repetition in the different chapters. Most of the material is well selected. However, some inconsistencies have unfortunately crept in. For example, on page 26, it is stated that calcium comprises 50 per cent of the weight of bone ash, while on page 84, 36 per cent is the figure used. The bibliographies on calcium and potassium are transposed. It seems unfortunate that, while tables are given of the mineral contents of foods, the standards of intake as compiled by the National Research Council in 1941 are not mentioned. Greater attention is paid to the pathological deficiencies than to ways in which the minerals function in positive nutrition, while the increased needs of the minerals for growth are scarcely mentioned.

In spite of these drawbacks, this is a readable little book from which the layman will gain an insight into the mineral metabolism of the body pleasantly and easily.

MARGARET L. FINCKE

OREGON STATE COLLEGE
CORVALLIS, OREGON

HANDBOOK OF COLLEGE ENTRANCE REQUIREMENTS. *William W. Hinckley*. Bulletin 1941, No. 13, Federal Security Agency, U. S. Office of Education. United States Government Printing Office, Washington, 1941. 79 pp. 14 × 23 cm. For sale by the Superintendent of Documents, Washington, 15 cents.

LABORATORY MANUAL TO ELEMENTARY GENERAL CHEMISTRY. *John C. Hogg*, M.A., M.C., Chairman, Science Department, The Philips Exeter Academy, and *Charles L. Bickel*, Ph.D., Instructor in Chemistry, The Philips Exeter Academy. D. Van Nostrand Company, Inc., New York, 1942. x + 283 pp. 51 figs. 17.7 × 24.3 cm. Looseleaf. $1.60.

Part I contains 32 excellently presented experiments, 44 per cent quantitative, dealing with the staples of elementary chemistry. These are graduated from explicit directions at the start to mere suggestions in Part II which contains 15 more advanced experiments, 85 per cent quantitative.

Designed to accompany "Elementary general chemistry" by the same authors, the experiments omit reference to colloids, organic compounds, and practical applications of chemistry. Even the classic experiment of soap-making is omitted. Emphasis is placed on good technics, the use of small quantities of materials, and accuracy. The directions are clear, concise, and amply illustrated. Prospective users should check the time available for laboratory work with the time required to perform the relatively large proportion of quantitative experiments. The emphasis placed upon the fundamentals of a sound training in unadorned chemistry is a welcome feature.

ELBERT C. WEAVER

BULKELEY HIGH SCHOOL
HARTFORD, CONNECTICUT

INDUSTRIAL CHEMISTRY OF COLLOIDAL AND AMORPHOUS MATERIALS. *Warren K. Lewis*, Ph.D., Professor in the Department of Chemical Engineering, Massachusetts Institute of Technology; *Lombard Squires*, S.M., Chemical Engineer with E. I. du Pont de Nemours Company, Wilmington, Delaware; and *Geoffrey Broughton*, Sc.D., A.I.C., Chemical Engineer with the Eastman Kodak Company, Rochester, N. Y. The Macmillan Company, New York, N. Y., 1942. xi + 540 pp. 205 figs. 15 × 23 cm. $5.50.

To all those interested in industrial uses of colloidal and amorphous materials this book will come as a most welcome contribution. While the book was written primarily to serve as a textbook it is filled with general information which will be of value to anyone interested in this field. The authors disclose that this book is the outgrowth of over thirty years of personal practical experience, that "it embodies the results of twenty-five years teaching the subject," and represents joint contributions of a large number of individuals who have been associated with the field of work at the Massachusetts Institute of Technology.

Approximately the first third of the book is devoted to the development of a scientific foundation upon which the theoretical treatment of the remainder dealing with industrial applications is based. An idea of the scope of the book can be gained from the following table of contents listed according to chapters: I. Structure of Liquids, II. Viscosity, III. Surface Tension, IV. Surface Tension and Orientation, V. Adsorption, VI. Suspensions, VII. Amorphous Solids, VIII. Emulsoids, IX. The Electrochemical Behavior of Colloids, X. Gelation, XI. Emulsions and Foams, XII. Crystalline and Amorphous States, XIII. Thermoplastics, Glass, XIV. Plasticization by Solution, XV. Paper, XVI. The Plastic Fibers, XVII. Leather, XVIII. Rubber, XIX. Ceramic Industries, XX. Synthetic Resins and Plastics, XXI. Textile Fibers.

The authors made no attempt to cover the entire field of important colloidal and amorphous materials but selected for treatment some of those of "major importance to industry or which seem likely to become important in the not too distant future." The materials selected for treatment are so diverse in properties that explanations of the phenomena relating to them made possible the application of nearly all of the more important principles of physical chemistry dealing with matter in the finely divided colloidal or amorphous state. The treatment of almost any chapter may at first seem too brief but a more careful reading will disclose that much information has been concisely presented. It is only by viewing the book as a whole that one realizes that nothing present could well have been omitted and that more extensive treatment would have made the book too large to be suitable as a text.

The reviewer believes that this very worthy book comes at an opportune time and that it will serve not only to emphasize the desirability of making available to students in chemical engineering and in chemistry the type of information presented, but that it will stimulate teachers to offer a course of this or of similar content.

F. E. BARTELL

UNIVERSITY OF MICHIGAN
ANN ARBOR, MICHIGAN

THE CHEMISTRY AND MANUFACTURE OF COSMETICS. *Maison G. de Navarre.* With a Chapter on the Federal Food, Drug and Cosmetic Act of 1938 by *Ralph J. Mill.* D. Van Nostrand Co., Inc., New York, 1941. xix + 745 pp. 92 tables and 179 figs. 15 × 23 cm. $8.00.

This large volume is by far the most ambitious and comprehensive work on cosmetics that has appeared to date in the English language. It can be used with profit by teachers, students, and even the most experienced of manufacturers.

In 34 chapters, the author gives an excellent presentation of the scientific and technical data necessary for the efficient compounding and production of various cosmetic preparations; brief résumés of the simple mathematics, chemistry, and other branches of science required; descriptions of raw materials, with special sections on certified colors, preservatives, anti-oxidants, and incompatibilities. Relatively few recipes are given, but those are guaranteed to be satisfactory if directions are carefully followed.

Special features of the book, not available elsewhere in this form, are the digests of U. S. and foreign patents on each type of preparation discussed; and a detailed exposition of the Federal Food, Drug and Cosmetic Act of 1938 (supplemented by helpful questions and answers). There are many good illustrations and tables, a voluminous bibliography, and a comprehensive index.

Even as the reviewer feels grateful for a work of such magnitude and general reliability, a few little motes must be viewed with disfavor. The classification of cosmetics in the introduction seems oddly arbitrary, labored, and questionable in spots. One wonders, for instance, at the omission of toilet soaps; and the miscellaneous groups deserve reshuffling. Hair dyes are conspicuous by their absence, although these equally valuable cosmetics could do with a good airing in some non-proprietary publications. The awe and studied care with which these important products are treated by many otherwise authoritative writers seems like just another manifestation of *Omne ignoto pro magnifico* (perhaps *terrifico*).

One could also wish that the bibliography were better organized, with the authors' names in alphabetic order. But this, as well as the matter of bulk, occasional inconsistencies in spelling and nomenclature, and other minor points should more properly be laid to lack of judicious editing and planning, rather than to any responsibility of the author.

The author has done an inestimable service both to industrial and to educational circles. Through his wide experience as a consultant, and his teaching at Wayne University he has learned what is useful to both groups, and his valuable book should be within easy reach of teachers and students of chemistry, and anyone else that is seriously interested in the background of this increasingly important field of scientific investigation.

FLORENCE E. WALL

NEW YORK UNIVERSITY
NEW YORK, N. Y.

THE "PARTICLES" OF MODERN PHYSICS. *J. D. Stranathan*, Ph. D., Professor of Physics, University of Kansas. First Edition. The Blakiston Company, Philadelphia, 1942. xvi + 571 pp. 218 figs. 15 × 23 cm. $4.00.

This is a textbook for a survey course in modern physics written for juniors, seniors, and graduate students. It is also an excellent reference book on the classical experiments in atomic physics as well as on the most recent work in the field. The "particles" include gaseous ions, the electron, positive rays, photons, alpha, beta, and gamma rays, atomic nuclei, positrons, neutrons, mesotrons, the neutrino, and cosmic rays. The Bohr-Sommerfeld picture of the atom is discussed briefly, and the application of wave mechanics is pointed out with no attempt being made to use the Schrödinger equation.

The author makes no claim to originality of material, but he has assembled most of the important experimental evidence for the existence of the many particles, and his theoretical discussion is simple yet thorough. Particularly happy is his selection of important and interesting nuclear disintegrations from the many available. The only adverse criticism which might be offered is against the common practice of physicists of giving the exact nature of the products of nuclear disintegrations without showing how the elements are identified. Errors are limited to a few mistakes in spelling and to the use in one equation of the symbol Ur.

Every teacher of general chemistry should know about these "particles"; probably few are able to keep up with the rapid advances in modern physics. This volume should enable the chemist to catch up with recent discoveries in this field. The author is well equipped for his task, and has succeeded in giving us an excellent and fairly complete résumé of the important work that has been published up to 1942.

ANDREW J. SCARLETT

DARTMOUTH COLLEGE
HANOVER, NEW HAMPSHIRE

FOOD ANALYSIS: TYPICAL METHODS AND THE INTERPRETATION OF RESULTS. *A. G. Woodman*, Associate Professor of Analytical Chemistry, Emeritus, Massachusetts Institute of Technology. Fourth Edition. McGraw-Hill Book Company, Inc., New York, N. Y., 1941. xii + 607 pp. 94 tables and 82 figs. 15 × 23 cm. $4.00.

This book is written primarily for the undergraduate student. It contains methods for the determination of the physical and chemical characteristics of certain typical groups of food materials. Quite a little space has been devoted to a discussion of the interpretation of results from analysis by use of these methods as they may indicate artificial manipulation of food products. The first chapter is devoted to general methods and equipment which may be used for analysis of all types of foods. Other chapters cover such topics as: The Microscopical Examination of Foods, Food Color and Preservatives, Milk, Cream and Ice Cream, Edible Fats and Oils, Carbohydrate Foods, Cocoa and Chocolate, Spices, Cider Vinegar, Flavoring Extracts, and Alcoholic Foods.

The content of this revised edition of the book is practically as originally arranged except for the omission of obsolete material, the revision of old methods to bring them up-to-date, and the addition of new methods. Because of the increased use of pasteurized milk, and also the fact that an attempt is often made to substitute heated milk for certified raw milk, methods have been added for the detection of heated milk. The section on food preservatives has been cut to a minimum and a list of dyes permitted under the new Food and Drugs Act has been added. New methods also have been introduced for the detection of the different forms of adulterations or substitutions which are commonly found in alcoholic beverages since prohibition was abolished.

The new edition, like the old, is well written, clear and concise. It is a very valuable book for the use for which it was intended, a text for undergraduate students who wish a good general foundation course in food analysis. While it does not give all of the methods of analysis which may be used, it does include the most reliable ones, and ones of general applicability to the different classes of foods. Selected references are given to more specific types of determinations.

BESSIE B. COOK

UNIVERSITY OF CALIFORNIA
DAVIS, CALIFORNIA

A COURSE OF INSTRUCTION IN THE QUALITATIVE CHEMICAL ANALYSIS OF INORGANIC SUBSTANCES. *Arthur A. Noyes*, late Professor of Chemistry, California Institute of Technology. Tenth Edition, Revised and Rewritten by *Ernest H. Swift*, Associate Professor of Analytical Chemistry, California Institute of Technology. The Macmillan Company, New York, N. Y., 1942. xv + 418 pp. 13 figs. 14 × 21.5 cm. $2.75.

The QUALITATIVE CHEMICAL ANALYSIS by A. A. Noyes has been held in high regard by chemists for more than thirty years; its friends will welcome the tenth edition, revised by Swift. One is impressed by the fact that its objectives, as stated thirty years ago, can be repeated in the current edition with so little revision. The objectives are: "...on the experimental side, to train the student of qualitative analysis in careful manipulation and exact method of procedure, such as are commonly employed in quantitative analysis....on the theoretical side, to make clear to the student the reason for each operation and result, and to accustom him to apply to them the laws of chemical equilibrium, and especially the principles relating to solubility and to ionization, complex-formation, and oxidation and reduction of substances in solution." The reviser has made some changes in arrangement and has revised and rewritten some experimental procedures, but these changes are in agreement with the stated objectives.

The contents are as follows:

Part I, entitled "The Course of Instruction," contains 10 pages of general directions and discussion, 25 pages of laboratory experiments, and 27 pages of questions. Part II on the system of analysis is subdivided into "procedures" each of which is introduced by discussion of the principles involved and followed by "notes" containing detailed explanations. In this part 236 pages are devoted to analysis for the basic constituents and 80 pages to analysis for the acidic constituents. This distribution of the theoretical discussion among the various procedures has the advantage of relating the theory directly with the experiment in which it is applied, but it has the accompanying disadvantage of a segmented presentation, for many of the principles apply to other procedures than the ones in which they are first introduced. The notes are excellent. They contain much detailed information and explanation, necessary for every serious student of qualitative analysis.

Swift favors macro methods. He suggests that centrifugation may be used to eliminate troublesome filtrations and that limiting the quantity of material decreases both the technical skill and the time required for analysis so that the number of analyses will compare favorably with that made by the semi-micro technic. Certainly he does not present the haphazard qualitative analysis which was so unsatisfactory that reaction from it has favored the recent trend toward semi-micro technics.

In our current debates on whether qualitative analysis should be taught in a separate course or in the course in General Chemistry, and on the macro versus semi-micro technic, we sometimes forget that the common objective is to train chemists so that they acquire the necessary theory and experimental skill early in their career. The curricular arrangement of the chemistry courses is secondary to their combined effect on students, and the methods used in separating ions is secondary to a knowledge of their behavior. From this point of view one may say that this is an excellent book which will fit into some curriculums and not into others. It can be recommended to all teachers who offer qualitative analysis with the objectives of the authors and who use macro methods. In addition it will be read profitably by all who are interested in the precise analytical separations and in the sensitivity of qualitative tests.

LUKE E. STEINER

OBERLIN COLLEGE
OBERLIN, OHIO

HOW TO SOLVE PROBLEMS IN QUANTITATIVE ANALYSIS. *Saul B. Arenson*, Professor of Inorganic Chemistry, University of Cincinnati. Thomas Y. Crowell Company, New York, N. Y., 1942. viii + 89 pp. 14.5 × 21.6 cm. $0.75.

A handbook on solving quantitative analysis arithmetic problems. The sections include computations, gravimetric analysis, normal solutions in neutralization, normal solutions in oxidation and reduction, normal solutions in precipitation titrations, and an appendix of logarithm and antilogarithm tables.

THE ELECTRON MICROSCOPE. *E. F. Burton*, Head of the Department of Physics, and Director of the McLennan Laboratory, University of Toronto; and *W. H. Kohl*, Development Engineer, Rogers Radio Tubes Limited, Toronto. Drawings by *Dorothy Stone*. Reinhold Publishing Corporation, New York, N. Y. 1942. 233 pp. 135 figs. 15 × 23 cm. $3.85.

In a very simple, clear, yet accurate style which emphasizes the fundamentals of the subject, the authors discuss the principles leading up to the development of the first electron microscope in America. The subjects discussed are: vision, light microscopes, nature of light, wave motion and wave motion media, the wave theory of light, the electromagnetic theory of light, the electron, the dual theory of light, the dual theory of the electron, the motion of electrons in electrical fields, electrostatic electron mirrors and lenses, magnetic lenses, the history of the electron microscope, the electrostatic electron microscope, applications of the electrostatic electron microscope, the compound electron microscope—magnetic type, and what the electron microscope can accomplish. A general but not exhaustive bibliography is given. The book is clearly and strikingly illustrated throughout. It may be recommended to any but the most technically informed reader.